DICTIONARY OF SIMPLE SUBSTANCES,

AND OF THEIR SIMPLE COMBINATIONS,

INDICATING THE TESTS BY WHICH THEY MAY BE IDENTIFIED.

ACETATES AND ACETIC ACID.

($C_4H_3O_3,HO$ or $\bar{A},HO$)

ACETIC acid is an acid which at a temperature under 60° Fahr. is crystallised and colourless, above that temperature it melts into a perfectly transparent liquid of a characteristic odour (that of vinegar); it has a burning, strongly acid flavour, and is, in that state, almost as corrosive as the most powerful mineral acids; it blisters the skin, and the sore thus produced is exceedingly painful. It dissolves in water, alcohol, and ether in all proportions, and boils at 248° Fahr.; its vapour is inflammable, and burns with a blue flame. It dissolves a great many substances, camphor, essential oils, fibrine, and it does not precipitate albumen.

Monohydrated sulphuric acid (oil of vitriol, SO_3,HO) and chlorine decompose acetic acid, two acids being the result, namely, *sulphacetic* and *chloracetic* acids.

Alcohol may interfere with the affinities and reactions of acetic acid so as to prevent its decomposing certain carbonates, or even its reddening blue litmus paper.

All acetates are decomposed by a red heat. A few of them, when so treated, yield acetic acid and a metallic residue; others yield *acetone;* and a carbonate of the base is left as residue.

All acetates are soluble in water, yet those of silver and of mercury are but sparingly soluble in that menstruum.

We must also add to the list of insoluble, or sparingly soluble acetates, those of tungsten, of molybdenum, the sub-acetate of lead ($6PbO,C_4H_3O_3$), which is insoluble in cold, and only sparingly soluble in hot water; the tribasic acetate of copper, $3CuO,2C_4,H_3O_3$, which is but little soluble in cold water, and which deposits a brown subacetate with excess of base, by boiling in water.

Solution of acetate of alumina (*unless the salt is perfectly pure*) becomes turbid by boiling, but it reassumes its original transparency on cooling.

Heated with dilute sulphuric acid, acetates are decomposed, and acetic acid, recognisable by its odour, is liberated.

Heated with a little caustic potash and arsenious acid to about 400° Fahr., an intolerably fetid and powerfully garlicky odour is evolved (acetyle of arsenic, or oxyde of cacodyle, C_4,H_6,As,O), which is very characteristic. The odour is very diffusible, and highly poisonous.

TESTS AND REACTIONS.

Fe_2Cl_3 . . .	*Nothing*. .	Solution of perchloride of iron produces nothing with free acetic acid; but if an excess of this reagent be poured into a solution of *neutral acetate*, or of acetic acid previously neutralised *exactly* with NH_3, no precipitate is produced, but the liquor becomes of a
	Dark red colour; .	if, however, an excess of NH_3 be poured into it, then a
	Reddish-brown .	precipitate is produced, which is Fe_2O_3. (See Table VIII., Observation *m*.)

ACETIC ACID.—ACETATES.

SO_3,HO + Alcohol.	*Odour of acetic ether.*	Acetates, heated with equal parts of concentrated sulphuric acid (SO_3, HO) and alcohol, evolve *acetic ether*, the agreeable odour of which is characteristic.
SO_3,HO . .	*Odour of vinegar.*	Acetates, treated by SO_3,HO, evolve acetic acid, recognisable by its odour.
AgO,NO_5 . .	*White* .	precipitate in neutral solutions; almost insoluble in cold, but more soluble in hot water.
Hg_2O,NO_5 . .	*White* or *yellowish*	. precipitate in neutral solutions, the precipitate is decomposed by hot water, mercury being revived so that the precipitate becomes *grey*.

The presence of acetic acid and of acetates is put beyond doubt by distilling with dilute SO_3, collecting the distillate, and digesting *in the cold* an excess of litharge in it; if acetic acid is present, the litharge will partly dissolve, and the solution will have *an alkaline reaction*. No other organic acid has this property. (See Table VIII., Observation *n*.)

The test with *caustic KO*, and AsO_3, so as to produce oxyde of cacodyle, alluded to at the beginning of this article, is likewise very characteristic.

ALUMINA.

(Al_2O_3)

Pure alumina is *white;* its hydrate is *white* also. When dried slowly, it becomes *yellowish* and horny, owing to the presence of organic matter, which it retains with great energy. After ignition it is still soluble in acids, but with the greatest difficulty. Its hydrate, however, is very readily soluble therein.

The salts of alumina are colourless, their taste is astringent, they all have an acid reaction, and are not precipitated by any acid.

TESTS AND REACTIONS.

NH_4S . . .	*White* . .	precipitate in neutral solutions. This precipitate of hydrate of alumina is soluble in KO. Its presence is rendered more apparent by boiling. (See Table V., Observation *f.* Table XVIII, Observations *d*—*i.*)
KO . . .	*White* . .	bulky precipitate of hydrate of alumina, immediately soluble in an excess of KO, even in the cold; a solu-
NH_4Cl	. .	tion of NH_4Cl, poured in this KO solution, reprecipitates the hydrate of alumina, especially with the help of heat, *and this is a distinctive character.* Remember, however, that phosphate of alumina behaves like pure alumina; wherefore the precipitate should always be examined for phosphoric acid. (See Phosphoric Acid.)
NH_3 . . .	*White* . .	precipitate of hydrate of alumina; which, according to Gay Lussac, is mixed with a subsalt, which cannot be decomposed by any excess of NH_3. This precipitate is soluble in a very large excess of NH_3, but is insoluble therein, if ammoniacal salts be present in any quantity.
		In very dilute solutions NH_3 produces no precipitate, since hydrate of alumina is slightly soluble in water.
KO,CO_2, or NaO,CO_2	. *White* . .	precipitate of hydrate of alumina insoluble in an excess of the reagent. The formation of this precipitate is accompanied by a disengagement of CO_2.
K_2Cfy . .	*White* . .	precipitate; but only after a time.
KO,SO_3 . .	*White* . .	*crystalline* precipitate of alum ($KO,SO_3 + Al_2O_3, 3SO_3$), especially after vigorous shaking.

Blowpipe.—Strongly heated upon charcoal, then moistened

with CoO,NO_5, and then strongly heated again, the mass assumes a ***beautiful blue colour.***

Several substances, and among others SiO_3, assume also a ***blue colour*** under this treatment, but none so intensely as Al_2O_3. (See Table XXXI., Observation *d.*)

The most characteristic tests for alumina are NH_3—NH_4S—$KO,+NH_4Cl$, and the blue colour produced with CoO,NO_5 before the blowpipe.

ALUMINIUM.

(Al.)

Aluminium is a metal of a white colour, with a bluish tinge, its hue being between that of platinum and zinc. Its specific gravity is 2·6; it is very sonorous, ductile, and it does not very sensibly oxydise by exposure to the air at the ordinary temperature; its fusing point is about that of cast iron. It is scarcely acted upon by water, but is soluble in the diluted acids, especially in HCl. It is also soluble in the solutions of potash or of soda, and is slightly magnetic.

AMMONIA.

(NH_3, or NH_4O.)

Ammonia is a gas of a pungent, penetrating odour, very soluble in water, to which it imparts its odour. This solution has a very caustic taste, and a powerful alkaline reaction upon test papers. Ammonia is expelled from its aqueous solutions by boiling. Most of its salts are colourless and soluble in water, and sublime or volatilise at a high temperature, with or without decomposition; they are isomorphous with the corresponding salts of potash, and have generally an acidulous reaction upon litmus paper. (See Table I.,—A, Observ. *t.*)

TESTS AND REACTIONS.

KO . . . . and other alkalies, triturated or boiled with salts of ammonia, evolve the characteristic

Odour of NH_3; but if in too small proportion to be thus detected, a glass rod, dipped in

	Thick white fumes.	moderately strong but not fuming HCl, or in strong $\bar{A}$, HO, immediately gives rise to the production of
$PtCl_2$ . . .	*Yellow* . .	precipitate of $NH_4Cl + PtCl_2$ similar to that produced by potash; but ignition leaves pure metallic platinum.
$\bar{T}$, 2HO	*White granular* . .	poured in excess and vigorously shaken produces a precipitate, similar to that yielded by potash, but sparingly, yet more soluble than tartrate of potash in water. It is also soluble in alkalies
HFl, $SiFl_2$. .	*White* . .	gelatinous precipitate.

The most characteristic tests for ammonia are the odour of NH_3, evolved by trituration with KO, or with CaO. (See Table XX., Observation *f*.)

ANTIMONY.

(Sb.)

Antimony is a white brittle metal, with a bluish tinge like zinc; its texture is crystalline and lamellar; its specific gravity is 6·702. It melts at about 800° Fahr. When allowed to cool down quietly after fusion, the surface of the melted mass has the aspect of the impression of fern leaves. Metallic antimony is not altered at the ordinary temperature by either atmospheric air or water. Nitric acid, even when diluted, transforms it into a white powder, antimoniate of protoxyde of antimony (SbO, Sb_2O_5), but it is not dissolved by that acid, in which respect it resembles tin. The other diluted acids have no action upon the metal. Hot concentrated SO_3, HO, in contact with it disengages SO_2, and produces $Sb_2O_3, 2SO_3$, which, when treated by water, yields a basic salt. HCl has scarcely any action upon it, but aqua regia dissolves it rapidly; when the aqua regia employed is in excess, it produces a perchloride Sb_2Cl_5, otherwise the result is a sesquichloride Sb_2Cl_3.

ANTIMONY AND SALTS OF ANTIMONY.

(Sb_2O_3.)

Sesquioxyde of antimony (Sb_2O_3) is in the state of white crystalline needles, or of a white powder. It fuses at a low red heat into a yellow mass, and crystallises into white needles on cooling. Exposed to a greater heat, it volatilises into a white powder. Heated in contact with the air, it becomes antimonious acid. It is soluble in HCl, and the solution is rendered milky by the addition of water; the milkiness is dissolved by $\bar{T}$,2HO.

The neutral salts of antimony are colourless, yellowish, or yellow; they all have an acid reaction on litmus paper, and are poisonous. They are generally decomposed by water, unless free muriatic or tartaric acids, or various organic substances be present, in which case no milkiness is produced by water. (See Table I.,—E, Observation *m*.) They are generally decomposed, at least partly, by a red heat; yet the chloride and bromide of antimony volatilise without decomposition.

TESTS AND REACTIONS.

HS . . . . *Orange-red* . precipitate (Sb_2S_3), in acid and in neutral solutions of sesquisalts, soluble in alkaline sulphurets, and in alkalies; also in HCl, with evolution of HS. If only a small quantity of antimony is present, or if, which is the same thing, the solution is very dilute, *and is neutral*, the liquid assumes only a red colour, but no precipitate is produced; if, however, some HCl is then added, the precipitate immediately takes place.

This precipitate may at once be identified as Sb_2S_3; because, if after having first dried it, it be dissolved in boiling HCl (by the action of which it is of course converted

into Sb_2Cl_3, with evolution of HS), the solution, on being poured into a long cylindrical glass full of water, will produce a dense white precipitate of $Sb_2Cl_3(2Sb_2O_3),HO$ (powder of Algaroth), which may be distinguished from the precipitate produced by salts of Bismuth by the action of NH_4S on the powder, the latter being turned black thereby, and also by the action of $\overline{\overline{T}},2HO$ in which the antimonial powder is soluble. (See also Table XVI., Observation *f*.)

Antimony may also be detected by hydrogen gas, as described in the Dictionary of Reagents. (See article Hydrogen.)

NH_4S . . .	*Red*, or *deep orange-red*	precipitate (Sb_2S_3), soluble in excess of the reagent; reprecipitated by an acid.
KO	*White* . .	*bulky precipitate* (Sb_2O_3,HO); soluble in a large excess of the reagent, especially with the help of heat.
NH_3 . .	*White* . .	*bulky precipitate* (Sb_2O_3,HO); insoluble in an excess of the reagent. If organic substances (such as $\overline{\overline{T}},2HO$, for example) are present, no precipitate at first is produced, but, after some time, a *white precipitate* falls down, which precipitate is insoluble in an excess of ammonia.
KO,CO_2, or NaO,CO_2, or NH_4O,CO_2	*White* . .	*bulky precipitate* (Sb_2O_3,HO), accompanied by a disengagement of carbonic acid.
K_2Cfy . .	*White precipitate*	in *concentrated*, but not in dilute, solutions, this precipitate, however, appears to be produced only by the water of the solution of the reagent.
Water		*Milkiness* re-dissolved by $\overline{\overline{T}},2HO$, and therefore the presence of that acid, or of other organic substances, and of excess of HCl, prevent this milkiness. (See Table IX., Observation *i*.)
(3K) 2Cfy .	*Nothing.*	

Tincture of Galls . .	*White* . .	precipitate ($Sb_2O_3, 3\overline{Q}t$).
Bars of iron or of zinc	. . . .	precipitate metallic antimony from its solutions, under the form of a black powder. (See Table XVI., Observation *e*.)

The most characteristic tests for antimony are, therefore, HS—the milkiness or precipitate produced by WATER in the solutions of the neutral salts, and the solubility of this precipitate or milkiness in $\bar{T}, 2HO$.

Blowpipe.—Mixed with cyanide of potassium, or with 1 part of NaO, CO_2, and 2 parts of KCy, and heated before the reducing flame of the blowpipe upon a charcoal support, bright and brittle metallic globules of antimony are obtained, which, being redissolved in HCl, and poured in water, produce a milkiness, which with the reaction produced by HS and NH_4S, is quite characteristic.

ARSENIC.

(As.)

Arsenic is generally in the form of grey—almost black—crystalline masses, which are easily pulverised. When newly volatilised, it resembles steel in colour, but it soon loses its lustre by exposure. It is insoluble in water, and has neither taste nor odour at the ordinary temperature; but exposed to a red heat, or thrown upon a piece of ignited charcoal, it is immediately volatilised, and a very strong, garlicky odour is evolved, *which is quite characteristic.* Heated in contact with the air, it volatilises in the shape of a white smoke, which is AsO_3. Its specific gravity is 5·75. Arsenic is not dissolved by HCl: fuming NO_5, or aqua regia, converts it into AsO_5. Treated by dilute NO_5, the solution contains AsO_3. Arsenic may be detected in solution by the tests for arsenious or arsenic acids, to which the reader is referred.

ARSENICAL COMPOUNDS.

ARSENIC ACID, ARSENIOUS ACID, SULPHURETS OF ARSENIC.

ARSENIATES AND ARSENIC ACID.

($As\,O_5$.)

Arsenic acid is a white deliquescent mass, in which state it is easily soluble in water. Anhydrous arsenic acid fuses at a low red heat into a white mass partly soluble in water, a residue being left which is AsO_3. Exposed to a higher temperature, it is volatilised, but is at the same time decomposed into $AsO_3 + O$.

Solution of SO_2 added to one of arsenic acid, or of the arseniates, reduces them to the state of arsenious acid and arsenites.

The neutral arseniates of alkalies are almost the only salts of that substance which are soluble in water. The neutral and basic arseniates can bear a strong heat without undergoing decomposition.

TESTS AND REACTIONS.

HS . . . . produces in *acid* solutions, after a somewhat long time, and occasionally after a very long time (twelve hours for example), a

Yellow . . . precipitate (AsS_5). (See Table IV., Observation *c*. Table XI., Observation *c*. Table XVI., Observation *c'*. Table XXIII., Observation *j*.) The formation of this precipitate is hastened by boiling the liquor immediately after saturation with HS, and letting it stand in a warm place The precipitate always appears first on the sides of the disengagement tube. In dilute solutions the precipitate takes a longer time. The precipitate is soluble in NH , and in NH_4S.

SO_2 . . . . Solution of sulphurous acid (SO_2), poured into a liquor containing AsO_5, converts it, with the help of heat, into AsO_3, with formation of SO_3,HO, thus:

$$AsO_5 + 2SO_2 = AsO + 2SO_3.$$

This conversion of AsO_5 into AsO_3 is often necessary, arsenious acid and arsenites being much more readily precipitated by HS than arsenic acid and arseniates. (See Table XXIII., Observations *j*, *k*.)

NH_4S . . . *No precipitate* is produced by this reagent in neutral, nor in alkaline solutions of arsenic acid and arseniates, because sulphuret of arsenic is soluble in NH_4S. If arsenic is present, however, the addition of NH_4S converts it into a sulphosalt, namely, sulpharsenio-sulphuret of ammonium (pentasulphuret of arsenic + sulphuret of ammonium, NH_4S, AS_5), which re-

\+ HCl . . . . mains in solution. If, however, an acid is now added, this double sulphuret is decomposed, and a *Yellow* precipitate of AsS_5 is thrown down, especially with the help of heat.

I have described under the head Arsenious Acid (AsO_3), a method by which the presence of As may be identified in sulphuret of arsenic.

AgO,NO_5 . . . . In neutral solutions,

Light reddish-brown . . precipitate ($3AgO$),AsO_5) very soluble in dilute NO_5, and in NH_3. (See Table VII, Observation *n*. Table XVI., Observation *c*.)

CuO,SO_3 + NH_3 . *Greenish-blue* precipitate ($2CuO$,) HO,AsO_5).

Before the blowpipe arseniates behave like arsenites. (See Arsenious Acid.)

The detection of arsenic acid and arseniates when mixed with organic matter, is the same as for arsenious acid. (See Arsenious Acid.)

The best tests for arsenic acid and arseniates are HS, and the treatment of the sulphuret obtained, so as to show the presence of arsenic therein.

The brown precipitate, produced by AgO,NO_5, is very characteristic. (See also Table XVI., Observation *e*. Table XXIII., Observation *m*.)

ARSENIOUS ACID AND ARSENITES.

(AsO_3)

The Arsenious acid of commerce is generally found either in white, opaque, porcelain-looking masses with a conchoidal fracture, or as a white powder. It is *almost tasteless,* and perfectly inodorous. It is volatilised by exposure to heat, but without decomposition, and yields an inodorous white sublimate.

Arsenious acid is but sparingly soluble in cold water (100 parts of water dissolve about 1 of AsO_3), but it is more soluble in hot water, and very soluble in HCl, and in solutions of the alkalies.

The alkaline arsenites alone are soluble in water, all the other arsenites are insoluble therein, but are decomposed by HCl.

TESTS AND REACTIONS.

HS		In *acid* solutions HS produces a
	Yellow . .	precipitate (AsS_3), soluble in NH_3 and in NH_4S, in KO, in KO,CO_2, and other alkalies; Insoluble in HCl; but decomposed by NO_5. (See Table IV., Observation *c*. Table XI, Observation *c*. Table XVI, Observation *c*.)
NH_4S . . .	*No precipitate* . .	is produced by this reagent in neutral, nor in alkaline solutions of arsenious acid, and of arsenites, because AsS_3

	is soluble in NH_4S. If As is present, however, the addition of NH_4S converts it into a sulphosalt (sulpharsenio-sulphuret of ammonium—that is, trisulphuret of arsenic + sulphuret of ammonium, NH_4S, AsS_3)
+ HCl	which remains in solution. If, however, an acid is now added, this double sulphuret is decomposed, and a
Yellow . . .	precipitate (AsS_3) is thrown down, especially with the help of heat.

Since, however, sulphuretted hydrogen produces a yellow precipitate in acid solutions of Cadmium, and of peroxide of tin, and although other reagents (such as AgO, NO_5 for example) are sufficient to distinguish arsenical solutions from every other, yet it is often important to prove the presence of arsenic at once in the yellow sulphuret produced by HS.

The following method will, I think, be found both expeditious and conclusive, since it resumes the best characteristics of that substance. Dry a portion of the yellow sulphuret of arsenic and mix it with 1 part of NaO, CO_2, and 2 parts of KCy, and introduce a portion of it into a small tube closed at one end, of about $\frac{1}{8}$ of an inch bore, and 5 or 6 inches long. This tube should be made of german-glass, that is, it should contain no lead, otherwise a metallic stain would be produced whilst heating, which stain would interfere with the experiment. This being done, expose the small portion in the tube to the heat of an ordinary spirit-lamp, and cautiously keep it in the flame for a few moments, that is, until the mass begins to char. The tube, during this part of the process, should be held quite horizontally, otherwise a little volatilised water sometimes condenses on the cold sides of the tube and may, by running down to the heated part, crack it, and of course spoil the experiment. When the little mass in the tube has turned black, heat it strongly by urging the flame of the spirit-lamp with a blowpipe, so as to fuse it. A metallic mirror of arsenic is immediately formed at a very short distance from the heated

point, which mirror can be shifted backwards and forwards by directing the flame of the blowpipe upon it. This volatility or shifting is one of the characteristics of the arsenical mirror.

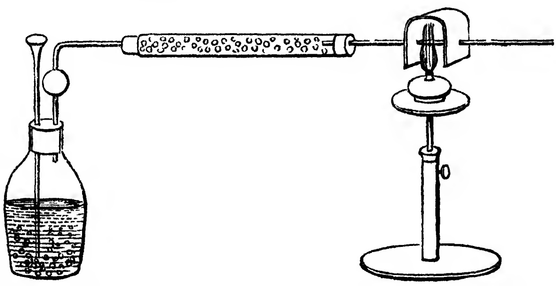
This done, cut the tube with a file close to the metallic mirror, and thrust the edge of the cut tube into the flame of the spirit-lamp ; an odour of garlic will then be perceived, which is another characteristic property of metallic arsenic, or of a very low degree of oxydisation of that substance.

Put now the part of the tube containing the mirror, and in fact the whole portion of the tube which has been operated upon, into a glass-beaker, and pour upon it a sufficient quantity of fuming NO_5 ; after a while dilute with water, and add an excess of AgO,NO_5, a white precipitate is produced which is cyanide of silver (AgCy), filter, and on adding carefully to the filtrate some NH_3, the characteristic light reddish-brown precipitate of arseniate of silver $(3AgO),AsO_5)$ is produced, which augments and becomes permanent when as much NH_3 has been added to the acid filtrate as will *nearly* neutralise it. Should the operator unguardedly have poured in more NH_3 than is necessary to neutralise the liquor, no precipitate will be formed, or that at first formed will be re-dissolved (arseniate of silver being immediately soluble in the slightest excess of ammonia, or of acid), but on neutralising the ammoniacal liquor with dilute NO_5, the precipitate is reproduced.

AgO,NO_5		This reagent produces, in neutral solutions, a
	Yellow . .	precipitate $(AgO)^2,AsO_5)$, very soluble in dilute NO_5, and in NH_3. If HCl is present at the same time, NO_5 should be added to the solution, then an excess of AgO,NO_5, that is to say, a greater quantity of AgO,NO_5 should be poured in than is necessary to precipitate the HCl, and the liquor, being violently shaken, is filtered. If, now, the filtrate be nearly neutralised with NH_3, the

		characteristic precipitate of arsenite of silver will be formed. (See Table VII., Observation *n*. Table XVI., Observation *c*)
CuO,SO_3 + NH_3	. . . .	In neutral solutions, this reagent produces a
	Bluish-green .	precipitate $(2CuO),AsO_3)$; a drop of NH_3 should be added.

When arsenious acid and the other compounds of arsenic are mixed with organic substances, it is absolutely necessary to destroy completely these substances, which otherwise would interfere with the production or the correct observation of the reactions by which those poisonous compounds may be detected. This is best done by heating the suspected mixture with about one-fourth of its weight of perfectly pure and concentrated SO_3,HO, by which the organic substances are converted into a brittle charcoal, which, after pulverisation, is then treated with NO_5, mixed with a small quantity of HCl, in order to convert the arsenic contained in the charcoal into arsenic acid (AsO_5). The mixture is then evaporated to dryness, digested in water, filtered, and introduced into a Marsh's apparatus represented in the following figure:—



A is a flask which must be large enough to contain the whole of the liquor to be operated upon, and yet not be more than three-quarters full. Its mouth must also be large enough to receive a cork perforated with two holes, through

one of which a funnel tube B is admitted, while the other is provided with a tube C bent at right angles, and connected with a larger tube D filled with asbestus and from the other extremity of which a tube of hard german-glass, containing no lead, E, issues, which tube E is provided with a metallic screen, under which a spirit-lamp is subsequently placed.

The apparatus being disposed as described, a few fragments of pure zinc are introduced into the flask or bottle A, water is then added, and a certain quantity of pure sulphuric acid. When all the air contained in the apparatus has been expelled by the hydrogen generated, *but not before*, the spirit-lamp is lighted so as to bring to a red heat that portion of the tube E placed between the metallic screen. The object of this preliminary operation is to make sure that the zinc and the sulphuric acid employed contain no arsenic. After having passed the stream of hydrogen through the heated tube for about half an hour, the tube, after cooling, should not be stained in the least, otherwise it would be a proof that the reagents employed are not pure. The asbestus placed into the large tube D is for the purpose of arresting any sulphate of zinc which might be mechanically carried by the hydrogen, and which, being reduced in the heated tube, might simulate a metallic mirror of arsenic.

Having thus ascertained that the reagents employed are pure, the suspected liquor is then admitted into bottle A, taking care to pour it slowly against the sides of the tube B so as not to convey air into the bottle, for a detonating mixture might then be formed which, being inflamed in coming in contact with the heated tube, might blow up the apparatus. If the suspected liquid contains arsenic, a metallic mirror is almost immediately produced in G. It is best also to add the suspected mixture only by very small portions at a time, otherwise the effervescence in the bottle may be so considerable and violent that a portion of the contents of the bottle may be forcibly ejected from the apparatus.

The metallic mirror thus obtained must now be tested to identify it as arsenic. This is done as follows :—

1. Examine whether the mirror can be shifted from place to place by heat.

2. Cut the tube with a file close to the metallic mirror, and, thrusting the edge of it in the flame, see whether an odour of garlic is produced.

3. Dissolve the metallic mirror in a few drops of fuming NO_5, dilute with water and filter, add AgO,NO_5, and then as much NH_3 as will nearly neutralise the liquor. If arsenic is present, a light brown precipitate will be produced, which is highly characteristic.

The testing of the mirror obtained is essential, since some other metals, and amongst them antimony, can combine also with hydrogen, and be reduced to the metallic state as is the case with arsenic ; but the mirror of antimony is not so volatile, and consequently cannot be shifted from place to place by heat so easily as arsenic ; and its treatment with NO_5 does not form a solution which can be precipitated by AgO,NO_5.

Arseniuretted hydrogen, being a highly deleterious gas, must never be allowed to be disengaged in the laboratory, but should be conveyed in the flue of the chimney.

Pure zinc may be prepared by M. Michelet's method, which is as follows :—

Melt some commercial zinc in an ordinary crucible, and when quite fluid pour the contents of the crucible into a deep pail full of cold water ; collect the zinc thus granulated, and if not in sufficiently small pieces, break the larger bits in a mortar. Put now the granulated zinc in a hessian crucible in alternate layers, with one-fourth of their weight of KO,NO_5, beginning with a layer of KO,NO_5, and ending with one of zinc. Heat the crucible, and after deflagration and fusion have taken place, remove the scories and run the zinc into an ingot mould.

REINSCH'S TEST.

Another most delicate process for the detection of arsenic consists in boiling the solid or liquid, supposed to contain arsenic, with about one-sixth of its weight of pure HCl., and introducing into the liquid a bar or slip of bright copper. Copper reduced by an electrotyping battery is best. After a short time, or immediately, if the quantity of arsenic is considerable, a grey or black deposit of metallic arsenic is deposited on the copper. The bar of copper coated with arsenic is then removed, washed with water, dried, and it is then heated in a reduction tube to obtain a metallic mirror, or the octahedral crystals of arsenious acid; or it may be put into a Marsh's apparatus, and treated therein as we said in speaking of that process. If the quantity of arsenic is large, or if the boiling be protracted, the arsenic generally peels off from the bar of copper; if small, on the contrary, the bar, instead of a black or steel-grey deposit, assumes only a bluish tinge. The copper should not be left too long in the acid, because, in that case, the action of the acid on the copper might simulate the presence of arsenic. The stain produced, however, under such circumstances, instead of having a metallic lustre, as when arsenic is present, has a dingy hue, and is easily removed by merely rubbing with the finger.

The operator should bear in mind that copper precipitates also Sb—Sn—Pb—Bi—Hg—Ag, but none of these deposits can be made to yield a white ring of arsenious acid in octahedral crystals by slowly heating the deposit in a reduction tube. In all cases the deposit, or sublimate, should be treated with a drop of fuming NO_5, the solution being now diluted with a little water, exactly neutralised by NH_3, and tested with perfectly neutral AgO,NO_5, will then yield a light brown precipitate if arsenic was present. (See Table XVI., Observation *e*.)

ARSENITES. (See ARSENIOUS ACID.)

BARIUM.

(Ba.)

Barium is a metal which has the lustre and colour of silver; it melts under a red heat, and is not volatilised by a higher temperature; it tarnishes by exposure to a damp atmosphere, and becomes coated with a white crust. Heated in the open air it burns with a red flame. Water is rapidly decomposed by it with a disengagement of hydrogen, and it is dissolved by the acids.

BARYTA—BARYTES.

(BaO.)

Pure oxyde of Barium (Baryta or Barytes) is greyish white, and combines greedily with water. This combination is attended with a great evolution of heat, and the result is hydrate of baryta (BaO,HO), which is soluble in 20 parts of cold water. Its concentrated solution deposits flattened hexagonal prisms ($BaO,10HO$). It is also soluble in alcohol and wood naphtha. It has a caustic and alkaline taste, and a powerful alkaline reaction upon coloured test-papers. It has a great affinity for carbonic acid, and must therefore be kept in well stoppered bottles. Its salts are colourless.

TESTS AND REACTIONS.

SO_3 . . . . or solutions of any soluble sulphate, produce immediately a
White . . . precipitate (BaO,SO_3), insoluble in acids and in alkalies.

$(2NaO)HO,PhO_5$ . . . . In neutral and in alkaline solutions,
White . . . precipitate $(2NaO),HO,PhO_5$, the formation of which is not interfered with by salts of ammonia.

NaO,CO_2, and other alkaline carbonates,
White . . . precipitate (BaO,CO_2), especially by

		boiling. Ammoniacal salts do not interfere with the formation of the precipitate. (See Table XXV., Observation *a*.)
KO	*White* . .	bulky precipitate (BaO,HO); in concentrated solutions, soluble in a great excess of water.
NH_3 . . .	*No precipitate,*	except the reagent contains NH_4O,CO_2, which frequently happens; in which case a
	White . . .	precipitate (BaO,CO_2), is of course produced. (See Table VI., Observation *b*.)
K_2Cfy . . .	*Nothing* .	in dilute solutions; but in concentrated solutions a
	White . .	*crystalline precipitate* appears after some time.
NaO,AsO_5 .	*White* . .	precipitate; insoluble in water; soluble in NO_5.
$SiFl_2,HFl$.	*White* . .	precipitate, which in moderately concentrated solutions, takes place immediately, but in dilute solutions only after some time. (See Table VI., Observation *e*.)

If hydrofluosilicic is not at hand, *succinate of ammonia* may be used to distinguish baryta from strontia, and from lime, because barytic solutions are precipitated by *neutral succinate of ammonia* in concentrated solutions, and after sometime in dilute ones. Whilst in dilute *strontia solutions* no precipitate is produced, and even in very concentrated ones some time is required before any turbidness appears; eventually, however, succinate of strontia is deposited in small crystalline grains very sparingly soluble in water. But in neutral solutions of lime, no precipitate whatever takes place, even after a long time, except the solution be very concentrated.

$KO,2CrO_3$.	*Yellow* . .	precipitate (BaO,CrO_3), soluble in an excess of acid. (See Table VI., Observation *f*; Table X., Observation *c*.)

Baryta and strontia behave in the same manner with most reagents; but the first is distinguished from the second by the following tests:

	Baryta.	*Strontia.*
$SiFl_2HFl$ . .	Immediate *white* precipitate,	*Nothing.*
$KO,2C_2O_3$.	Immediate *yellow* precipitate;	*Nothing* at first, yellow precipitate after some time.
Alcohol . .	*Nothing;*	red flame.

The most characteristic tests for Barium are SO_3,HO, or a soluble sulphate, and $SiFl_2HFl$.

BENZOATES AND BENZOIC ACID.

($C_{14}H_5O_3,HO$ or $\overline{Bz},HO$.)

Benzoic acid is white, and crystallises in hexagonal needles, or in the form of a crystalline powder, that obtained by sublimation is in flexible scales of a pearly whiteness. Pure benzoic acid has no smell, but that which has been obtained from benzoin has always a balsamic odour. The hydrated acid reddens litmus feebly; heated in the air it volatilises, and its fumes are very acid and irritating. (See Table XXII., —A, Observation *d.*) It is soluble in about 200 parts of cold, in 25 of hot water, and in two parts both of alcohol and of ether. When its aqueous solution is heated, it volatilises with the steam.

The benzoates of alkalies, and that of manganese are very soluble in water, and crystallise with difficulty. Benzoate of lime requires about 20 parts of water for its solution, those of lead, baryta, and strontia are very sparingly soluble in cold water; that of ammonia is deliquescent.—All the soluble benzoates have a fresh salty taste, and are decomposed by most acids, benzoic acid being thereby separated in the state

of a white precipitate, provided the solution is sufficiently concentrated. All acids behave with the insoluble benzoates in the same manner, provided the base of the benzoate be such as to form a soluble salt with the acid employed.

TESTS AND REACTIONS.

Fe_2Cl_3 . .	*Pale reddish-brown* . .	and voluminous precipitate ($Fe_2O_3,3\overline{Bz}$) immediately decomposed by NH_3, and acids; wherefore the solution tested by Fe_2Cl_3 must be perfectly neutral. (See Table VIII., Observations *j, k, l*)
$PbO,\overline{A}$ . .	*No precipitate*	in solutions of free benzoic acid, or of benzoate of ammonia, at least immediately, but in solutions of benzoates of the fixed alkalies a
	White . . .	flaky precipitate is produced.
Alcohol + NH_3 + BaCl	*No precipitate*	in solutions of free benzoic acid, nor of the benzoates of alkalies.
HCl		Hydrochloric (and other strong acids) produces a
	White . . .	precipitate of benzoic acid, in solutions of benzoates, as was said at the beginning of this article.

The most characteristic tests for Benzoic acid are its volatility and its reaction with Fe_2Cl_3.

BISMUTH.

Metallic bismuth is brittle, easily pulverised, and of a reddish white colour; it has a lamellar structure, and crystallises easily in beautiful cubes, or in pyramids derived from the cube and generally covered with a thin film of oxide, so thin as to be iridescent. Its Sp. gr. is 9·8 or 9·9. It melts at about 500° Fahr., and volatilises at a very high temperature. Dry air has no action on bismuth at the ordinary temperature,

but it is slightly oxydised in a damp atmosphere. Heated, however, in the air it becomes speedily converted into oxyde of bismuth (BiO). Dilute HCl, or SO_3, have only a very slow action on this metal; concentrated SO_3,HO dissolves it with the help of heat under disengagement of SO_2, but NO_5, even diluted, and aqua-regia dissolve it with great rapidity. Strong or fuming NO_5 has so violent an action, that the metal sometimes becomes red hot when such an acid is poured upon it.

OXYDES AND SALTS OF BISMUTH.

Oxyde of bismuth (BiO) is the base of all the salts of this metal; it is yellow, it becomes of a deeper hue when heated, but returns to its original colour on cooling. It is fusible at a high temperature, and in that state it fluxes with silica. Hydrated oxyde of bismuth is white. The salts of bismuth have all an acid reaction, and they are decomposed by water into a subsalt which is precipitated, whilst an acid salt remains in solution; if, however, an excess of acid be present, this decomposition does not take place.

TESTS AND REACTIONS.

HO . . . . *Milkiness*, or *White* . . } precipitate ($BiO, NO_5 + 3BiO, HO$); *insoluble in* $\bar{T}$,2HO. If too much acid is present, the milkiness is not produced. The excess of acid should, therefore, be neutralised beforehand, or else a large quantity of water must be employed. (See Table I.,—E., Observation *m*; Table IX., Observation *i*.)

The presence of organic substances does not interfere with the milkiness produced by water.

HS . . . . *Black* . . precipitate (BiS); but if only a small quantity is present, a *Brown colour* is produced.

If the liquid contains too much HCl, or, in fact, any other

mineral acid in excess, no precipitate will be produced by HS, unless the solution be much diluted with water, or unless the great excess of acid be neutralised by NaO,CO_2, or better still, by NH_3.

NH_4S . . .	*Black* . .	precipitate (BiS); insoluble in excess.
KO, or NH_3 . .	*White* . .	precipitate (BiO, HO); insoluble in excess, and becoming yellowish by boiling.
KO,CO_2		(or other alkaline carbonates),
	White . .	precipitate (BiO,CO_2); insoluble in excess.
K_2Cfy . . .	*White* . .	precipitate (Bi_2,Cfy); insoluble in HCl.
$K_3,3Cfy$ . .	*Dingy-yellow*	precipitate ($Bi_3,2Cfy$); soluble in HCl.
Infusion of Galls	*Orange-yellow*	precipitate.
KO,CrO_3 . .	*Yellow* . .	precipitate (BiO,CrO_3); *soluble in* NO_5, and *insoluble in* KO. (See Table IV., Observation *q*)

A bar of zinc, of copper, or of tin, precipitates bismuth from its solutions in the state of a black spongy mass.

Mixed with NaO,CO_2, and heated in the inner flame of the blowpipe upon a charcoal support, a very brittle metallic bead is obtained, accompanied with an orange incrustation. (See Table I.,—B, col. 16.)

The most characteristic tests for bismuth are HS—HO—KO,CrO_3, and the blowpipe.

BORATES AND BORACIC ACID.

(BO_3.)

Anhydrous boracic acid is a colourless, brittle glass, which becomes opaque in course of time, and is somewhat difficultly soluble in water. Hydrated boracic acid forms lamellar;

colourless, and inodorous crystals, requiring about 20 or 25 parts of cold water for their solution, and that solution imparts only a vinous red colour to tincture of litmus. It is more soluble in boiling water. It is a very weak acid, and may be eliminated from all its combinations by most other acids; yet on account of its *fixity*, it decomposes the salts of the most powerful acids when exposed with them to a high temperature. Boracic acid, in a state of fusion, dissolves a great number of metallic oxydes, forming with them vitreous beads, the colour of which is one of the means by which these oxydes may be identified. Boracic acid is soluble in alcohol, and when its solution in that menstruum or in water is boiled, it volatilises with the steam or vapours of these liquids. (See Table XXII.,—A, Observation *b*.) Yet the acid so volatilised can be easily collected in a receiver in the state of shining spangles. Crystallised boracic acid contains three equivalents, or 35 per cent. of water ($BO_3,3HO$).

The borates of alkalies are soluble in water; the other borates are insoluble, or only sparingly soluble therein; but the borates of the earths dissolve freely in liquids which contain a sufficient quantity of chlorides of earths. In the presence of water they are decomposed by sulphuric, muriatic, and nitric acids, boracic acid being separated, which imparts to alcohol the property of burning with a fine golden-green flame. (See Table XIV., Observation *a*, § 2.) The solutions of borates of alkalies have an alkaline reaction on litmus and on turmeric papers. Borates of earths are precipitated by NH_3 from solutions which do not contain a sufficient quantity of ammoniacal salts. (See Table XXII.,—A, Observation *e*.)

TESTS AND REACTIONS.

BaCl . . . *White* . . precipitate (BaO,BO_3) in moderately concentrated solutions, soluble in acids, and in salts of ammonia; consequently, no precipitate is produced by that reagent in solutions

		which may contain much ammoniacal salt. (See Table I.,—D, Observation *d.*) Borate of baryta is also completely soluble in a large quantity of water. (See Table VII., Observations *q*, *b'*. Table XXI., Observation *c*, *j* Table XXII.,—A., Observation *h*.)
CaCl . . .	*White* . .	precipitate (CaO,BO_3); soluble in acids, and in NH_3. (See Table VII., Observation *v*.)
AgO,NO_5 .	. . .	In concentrated solutions,
	White . . .	precipitate (AgO,BO_3); soluble in dilute NO_5, and in NH_3. (See Table VII., Observation *c'*.)
		In dilute solutions,
	Brownish . .	precipitate, which is AgO.
HgO_2,NO_5 . .	*Light dingy-brown* .	precipitate, soluble in NO_5.
SO_3 + Alcohol	. .	Concentrated SO_3,HO added to a solution of a borate, produces a precipitate in the form of
	White spangles . .	. (BO_3).
		If a borate, previously reduced to powder be drenched with concentrated SO_3,HO and alcohol being poured upon it, be then inflamed, it will burn with a
	Greenish yellow flame.	(See Table I.,—C, Observation *c*. Table XXVI.,—A, Observation *c*.)

When chlorides are present, the characteristic colour of the flame is somewhat interfered with, because hydrochloric ether is then formed, which burns with a blue flame.

The borates which have been mixed with SO_3,HO or which retain still a small quantity of that acid, should never be exposed to a red heat in platinum crucibles, for although sulphuric acid, and boracic have *separately* no action upon platinum, they on the contrary, when combined, attack platinum powerfully.

If a borate be pulverised with fluor-spar, and then heated with a weight of concentrated sulphuric acid (SO_3,HO) several times that of the mixture, a gas is disengaged (fluoride of boron—BFl_3) in the form of easily recognisable thick white fumes, as they come in contact with the air, which gas has the property of charring paper.

The principal tests for BO_3 and the borates, is the greenish-yellow flame which they impart to the flame of alcohol after treatment with SO_3,HO. Boracic acid alone gives that colour to the flame of alcohol, the borates do not behave so, unless previously decomposed by an acid.

The following test recommended by Turner is also valuable: it consists in reducing into very fine powder, the mineral or compound in which a borate is suspected, with about 4½ parts its weight of bisulphate of potash, and 1 of fluorspar; the mixture well pulverised is then kneaded into a paste with a little water, and a pellet of the said mixture, on being heated before the blue or reducing flame of the blowpipe on a loop of platinum wire, will impart a beautiful and pure green colour to the oxidising or outer flame of the blowpipe.

BORON.

Boron is a dull, greenish-brown powder, heavier than water, which burns in the air when heated, producing boracic acid (BO_3). It is readily attacked, with the help of heat, by nitric acid, and aqua-regia, by which it is converted into BO_3, Chlorine gas converts it into chloride of boron. Boron is closely allied to silicon, and is the basis of boracic acid.

BROMATES AND BROMIC ACID.

(BrO_5)

Bromic acid has many properties in common with chloric acid. Like the latter, it does not exist in the anhydrous state;

it is only feebly acid; it reddens litmus paper at first, but afterwards it decolorises it; it is decomposed by sulphurous acid, phosphorous acid, sulphuretted hydrogen, and by all the hydracids. It is decomposed also by sulphuric acid, which takes up its water, oxygen and bromine being at the same time disengaged. Bromates on being heated evolve oxygen, and are converted into bromides. Nearly all bromates are soluble in water. When thrown upon ignited charcoal they deflagrate.

TESTS AND REACTIONS

BaCl . . .	*White* . .	precipitate, soluble in a large quantity of water (BaO, BrO_5).
AgO,NO_5 .	*White* .	precipitate, (AgO,BrO_5), soluble in NH_3 and in NO_5, but with great difficulty. It is distinguished from chloride of silver by its being only slightly blackened by daylight, by its deflagrating when thrown upon ignited charcoal, and by its action with SO_3, HO (*see below*) (See also Table VII., cols. 22, 23.)
$PbO, \bar{A}$ . .	*White* . .	precipitate (PbO,BrO_5), soluble in a large quantity of water.
SO_3, HO . . .	*Hyacinth red colour* . .	treated in a test-tube, with cold concentrated SO_3,HO, fumes of bromine, of a hyacinth red colour are produced, oxygen being also evolved at the same time. The same effect is produced by heating the suspected substance in a glass test-tube, with about twice its bulk of bisulphate of potash.

The most characteristic tests for Bromic acid and Bromates are the white precipitate produced by AgO,NO_5, but more particularly the hyacinth-red fumes disengaged by treatment with cold SO_3,HO. When exposed to a red heat, bromates,

being converted into bromides, afford then the distinctive reactions of HBr and of bromides (*which see*).

The distinctive characters between bromides and bromates have been alluded to. (See Table XI., Observation *i.*)

BROMIDES. (See Hydrobromic Acid.)

BROMINE.

At the ordinary temperature, bromine is a liquid of an intense hyacinth or brownish-red colour, very volatile; the fumes are also of a hyacinth-red colour. It is very poisonous; its odour is very strong, suffocating, resembling chlorine or iodine. Its specific gravity is 2·966, that of its vapour is 5·393. It is slightly soluble in water, more so in alcohol, and in all proportions in ether. The aqueous solution has bleaching properties. A solution of KO at once destroys the brownish red colour of a solution of bromine, and transforms it into bromide of potassium and bromate of potash.

CADMIUM.

Cadmium has very nearly the same colour as tin, but is a little harder. It is malleable and ductile; it melts below 500°, and is almost as volatile as mercury; its specific gravity is 8·604, or when laminated 8·69. At the ordinary temperature it is not much oxydised by exposure even to a moist atmosphere. Heated in the air, it burns and emits brownish fumes of CdO. HCl, moderately strong, dissolves it rapidly, especially with the help of heat, and hydrogen gas is evolved. Sulphuric and acetic acids behave in the same manner, but with the latter acid the action is slow. Nitric acid is the best solvent of this metal.

Aqueous solution of SO_2 attacks cadmium rapidly also, and the result of the action is, according to Messrs. Fordos

and Gélis, a mixture of sulphate and of sulphuret of cadmium.

OXYDE AND SALTS OF CADMIUM.

(CdO.)

Dry oxyde of cadmium is a dark-brown, light-brown, or black infusible powder, which has sometimes a crystalline aspect.

Hydrate of oxyde of cadmium (CdO,HO) is white.

Both oxydes dissolve freely in acids, and in caustic ammonia. Its neutral salts redden litmus paper. Its salts are colourless, and have a disagreeable metallic taste.

TESTS AND REACTIONS.

HS and NH_4S	*Fine yellow*	precipitate (CdS) in acid and in neutral solutions, this precipitate is insoluble in an excess of NH_4S, which distinguishes it perfectly from the yellow precipitate produced by HS in solutions of antimony, of arsenic, and of peroxyde of tin, which are soluble in an excess of NH_4S. The solutions of Cd when very acid, are not precipitated by HS, unless previously diluted with a pretty large quantity of water.
KO	*White*	precipitate (CdO,HO), insoluble in an excess.
NH_3	*White*	precipitate (CdO,HO); soluble in an excess.
KO,CO_2 and other alkaline carbonates,	*White*	precipitate (CdO,CO_2), insoluble in excess. Ammoniacal salts do not interfere with the formation of this precipitate. (See Table XVII., Observation *e*.)

Blowpipe.—Mixed with NaO,CO_2, and heated in the reducing flame of the blowpipe upon a charcoal support, no metallic bead is obtained, but a brown incrustation is formed,

and brown fumes are evolved from the mass. (See Table I., —B, Observation *l*.)

CALCIUM.

(Ca.)

Calcium is a silver-white metal, which speedily becomes converted into lime by exposure; it is obtained only with difficulty, and is therefore not likely to be met with for analysis. Protoxyde of calcium (CaO), or pure lime, has many of the properties of baryta and of strontia, but in a less marked degree. It is white, and is generally met with in masses possessing considerable hardness. It has a caustic alkaline taste, combines powerfully with water, with evolution of heat, the result being slaked lime (hydrate of lime, CaO, HO), which is slightly soluble in water, and *more so in cold than in hot water,* the solution absorbing rapidly the carbonic acid of the air, so that the lime is converted into CaO,CO_2, which falls down. The salts of lime are colourless, and have a saline, bitter taste.

TESTS AND REACTIONS.

SO_3 (dilute)		and the soluble sulphates, poured in very concentrated solutions, produce a
	White . . .	precipitate (CaO,SO_3), completely soluble in acids, and in a large quantity of water. No precipitate takes place in dilute solutions
$(2NaO),HO,PhO_5$ . . .		in neutral and in alkaline solutions
	White . . .	precipitate,
NaO,CO_2		and other alkaline carbonates,
	White . . .	precipitate (CaO,CO_2) (See Table XXV., Observation *a*.)
KO	*White* . . .	gelatinous precipitate.
NaO . . .	*White* . . .	gelatinous precipitate.
$SiFl_2,HFl$.	*No precipitate.*	
$KO,2CrO_3$.	*No precipitate.*	

$NH_4O,\bar{O}$ . . *White* . . . precipitate ($CaO,\bar{O}$); insoluble in water, and in acetic acid; but very soluble in NO_5, in HCl, and slightly so in $\bar{O},HO$; hence the necessity of using $NH_4O,\bar{O}$ instead of $\bar{O},HO$, or of adding a little NH_3, which, under all circumstances, considerably exalts the delicacy of the test. (See Table VI., Observations *j*, *k*.)

The insolubility of the white precipitate of oxalate of lime in water and in acetic acid, and its ready solubility in NO_5 and in HCl, is a characteristic reaction.

Alcohol . . . . added to the solution, or mixed with the pulverised salt of lime, and inflamed,

Yellowish carmine flame.

The most characteristic tests for the salts of calcium are $NH_4O,\bar{O}$ and SO_3,HO + alcohol.

CARBON.

Diamond is pure carbon. Graphites or plumbago, however, is carbon in an almost pure state also. All the forms of carbon, except diamond, are black, and carbon in whatever state is infusible. When heated sufficiently high, it combines directly and energetically with oxygen, producing CO and CO_2. Hydrogen, phosphorus, arsenic, nitrogen, boron, silicium, chlorine, bromine, and iodine, have no direct action on carbon. Most varieties of carbon are oxydised by boiling in NO_5, producing carbonic acid and artificial tannin, soluble in an excess of acid, which is tinged brown by it. Chlorine gas and solution of pure potash have no action on carbon, but if fused with an alkaline carbonate, it becomes converted into CO_2, which is disengaged, and the residue consists of the pure alkali, alone or mixed with charcoal, if a sufficient quantity of carbonated alkali has not been used. Heated with KO,NO_5, deflagration is produced.

CARBONATES AND CARBONIC ACID.

Carbonic acid at the ordinary temperature and pressure is a colourless, incombustible, and irrespirable gas, heavier than atmospheric air. It has a sour taste; it is soluble in its own bulk of water, and the solution has an acidulous taste; it imparts to litmus paper a vinous red colour, but the blue colour reappears merely by exposure, or more rapidly by applying heat. Boiling altogether expels carbonic acid from its solutions. All carbonates, except those of the alkalies, are decomposed by heat, especially if aqueous vapour (steam) is passed through at the same time.

The carbonates of alkalies are soluble in water, all other neutral carbonates are insoluble therein. Carbonates of lime and of baryta are soluble in water containing an excess of carbonic acid.

Carbonates are decomposed by all the acids which are soluble in water, and the decomposition is accompanied by an effervescence, *provided an excess of acid be employed,* and provided also the solution (if operating on one) is not too dilute, otherwise no effervescence will take place, because, in that case, the carbonic acid either remains in solution, or combines with the neutral carbonate so as to convert it into a bicarbonate. (See Table I.—E. Observations *e, f.*) In order to decompose carbonates by an acid, it is generally necessary to apply heat, and also to dilute the acid with water, since several carbonates are not decomposed in the cold by concentrated acids. Such is more especially the case with MgO,CO_2,—BaO,CO_2,—PbO,CO_2, and FeO,CO_2. Dolomite (magnesian limestone) produces only a very sluggish effervescence, even when treated by dilute HCl, unless heat is applied. We must also except two acids which have not the power to

produce an effervescence with carbonates; they are hydrocyanic and hydrosulphuric acids.

The neutral carbonates which are *soluble in water*, have a strong alkaline reaction on test papers ; the acid carbonates have also an alkaline reaction, but less powerful than that of the neutral carbonates.

Most metallic solutions are precipitated by those of the carbonates, which are soluble in water.

TESTS AND REACTIONS.

$CaO+aq$.		(Lime-water) produces a
	White . .	precipitate (CaO,CO_2), insoluble in water, soluble, with effervescence, in HCl. An excess of the reagent must be employed, since bicarbonate of lime ($CaO,2CO_2$) is soluble in water. This, in fact, is the reagent usually employed to identify the presence of carbonic acid. The substance under examination is put into a large test-tube, or small flask, and an excess of dilute hydrochloric acid being poured upon it, the flask is then closed with a perforated cork, provided with a tube, plunging in a test-glass or beaker, containing lime-water. The formation of a milkiness, or white precipitate, (CaO, CO_2), insoluble in water, soluble with effervescence in HCl, indicates the presence of carbonic acid. (See Table I—E, Observation *g;* Table XXVII.—A, Observation *c*)
CaCl, or BaCl. .	*White* . .	precipitate (CaO,CO_2, or BaO,CO_2) ; in solutions of neutral carbonates; in solutions of bicarbonates, no precipitate is produced by these reagents, except by boiling, since bicarbonate of lime and of barytes are soluble in water.

The best test for carbonic acid, is the *white precipitate* produced by passing the gas evolved by treatment with an acid through lime water, and the *effervescence* produced by pouring an acid upon this white precipitate.

CERIUM.

Cerium is a greyish powder, of a very refractory nature, and which is not volatilisable by heat. It is not quite so hard as cast-iron; it decomposes water slowly at the ordinary temperature, hydrogen being disengaged, but the decomposition of boiling water by this metal is rapid. Heated with chlorate or with nitrate of potash it *detonates.* The oxyde of cerium, obtained by precipitating the double sulphate of cerium and potash directly derived from cerite* by KO,CO_2 contains the oxydes of two other metals, lanthanum and didymium. The oxydes of these three metals behave with reagents in the same manner.

TESTS AND REACTIONS

KO, or NaO, or NH_4O.	*White*	bulky precipitate, insoluble in excess.
NH_4S.	*Dingy-white*	precipitate, insoluble in excess.
HS.	*Nothing*	
KO,CO_2, or NaO,CO_2	*White*	precipitate, insoluble in excess. (See Table XXXI. Observation *e.*)
$\overline{O}$, HO.	*White*	precipitate, insoluble in $\overline{A}$, HO.
KO,SO_3.	*White*	crystalline precipitate; almost insoluble in water.

CHLORATES. (*See* CHLORIC ACID.)

CHLORIDES. (*See* HYDROCHLORIC ACID.)

* Cerite is a hydrated silicate of peroxide of cerium, mixed with a little carbonate of lime. This mineral is found in the copper mine of Bastnaes, near Riddarhyttan, in Sweden.

CHLORATES—CHLORIC ACID.

(ClO_5,HO)

Chloric acid is a strongly acid, colourless, odourless, oily liquid, soluble in water in all proportions, and which cannot be obtained in the anhydrous state. It is decomposed by sulphurous acid, the result being hydrochloric and sulphuric acids.

Hydrochloric acid, mixed with chloric acid, or a chlorate, forms a kind of aqua regia, chlorine, or a mixture of chlorine and of hypochloric acid being disengaged. Sulphuretted hydrogen decomposes chloric acid, chlorine, a little sulphuric acid, and a deposit of sulphur being the result. Chloric acid undergoes spontaneous decomposition, and is a powerful oxydising agent.

All chlorates are soluble in water, cholrate of potash being the least soluble of all these salts. The chlorates of alkalies, and in fact all chlorates are decomposed at a red heat. The chlorates of alkalies, when so treated, part with their oxygen, and a chloride of alkali remains; most other chlorates evolve oxygen and chlorine, a metallic oxide being left as a residue. Heated with charcoal, or other combustible substances, chlorates deflagrate smartly. mixed with phosphorus or sulphur, and, struck on an anvil, or other hard substance they detonate in a most violent manner; mixed with sugar, and moistened with a drop of concentrated SO_3,HO, they burst energetically into flame.

TESTS AND REACTIONS

Solution of Indigo + SO_3,HO or SO_2. . . . . . If a solution of sulphate of indigo is added to the liquor under examination, in sufficient quantity to impart thereto a slight blue colour, then some concentrated SO_3,HO, and the whole be boiled, the

Blue colour disappears exactly as is the case when nitrates are treated in the same manner. If

	instead of sulphuric acid, sulphurous acid is employed, the effect will be the same.
SO_3,HO	Concentrated SO_3,HO, poured into a concentrated solution of a chlorate, or upon a solid chlorate, produces an
Intense yellow colour . . .	accompanied by an evolution of a gas of a
Greenish colour	(ClO_4), and of a peculiar odour. Heat must be avoided, for fear of an explosion. Small quantities only should be used. (See Table VII., Observation *v*,.)
KCy	Mixed with cyanide of potassium, and heated on a piece of platinum foil, or in a test-tube, they deflagrate, and detonate violently. Only a small quantity should be used in the experiment.
Microcosmic salt + brass wire.	Heated with a bead of microscosmic salt upon a brass wire, over the smallest possible flame of a spirit-lamp, *a beautiful blue colour* will be imparted to the flame. This reaction is common to all substances containing chlorine.

AgO, NO_5 produces no precipitate in solutions of chlorates, because chlorate of silver (AgO, ClO_5) is soluble, but after ignition or calcination, the chlorate being converted into a chloride, its solution will then produce a precipitate in that of AgO, NO_5.

The most characteristic tests are the *yellow colour, and the odour of the gas* evolved by treatment with concentrated SO_3,HO, and the *blue flame* imparted to the flame of a spirit lamp when a particle is heated with microcosmic salt on a brass wire.

CHLORINE.

Chlorine is a gas of a yellowish-green colour, of an intensely suffocating odour, producing violent coughing when inhaled in even small quantities. Its specific gravity is 2·47, consequently, 100 cubic inches weigh 76·6 grains. It is soluble to a considerable extent in cold water, but this solubility diminishes rapidly as the temperature increases; it diminishes also if the temperature sinks below a certain point, the maximum of solubility being at about 47° Fahr., at which temperature 1 volume of water can take up about 3 volumes of gas, whilst at 32° Fahr. the water can only take about 1·5 its volume of the gas. Chlorine cannot support combustion, yet several substances, in a pulverised state, such as antimony and arsenic, burn vividly when poured into a jar full of the gas.

Chlorine, in presence of moisture, or in solution, possesses energetic bleaching properties. The solution has the colour of the gas, and becomes gradually and spontaneously converted into HCl, oxygen being disengaged; but when left exposed to the direct rays of the sun, ClO is produced, which, however, is soon decomposed by the HCl which is formed at the same time in the solution; according to M. Barreswill a little ClO_7 is also produced.

Chlorine and hydrogen have a very great affinity for each other, provided direct or diffuse light be admitted; for, in the dark, and at the ordinary temperature, these two gases do not react on each other. The combination of hydrogen and of chlorine, under the direct rays of light, is accompanied by a violent detonation; but in diffuse daylight it takes place quietly; in the dark, or in yellow, green, or red coloured glass vessels, no combination takes place. Mercury and most other metals absorb chlorine, and are converted into chlorides; chlorine gas cannot therefore be collected over mercury. The solutions of the fixed alkalies absorb it also, and become con-

verted into chlorides and chlorates. Carbonated alkalies absorb it also, but then carbonic acid is disengaged, and a hypochlorite is produced instead of a chlorate. Nitrate of silver is converted by chlorine into chloride of silver. Ammonia absorbs chlorine, and nitrogen is disengaged. Solutions of neutral ammoniacal salts produce chloride of nitrogen, an oleaginous looking liquid, which is fearfully explosive.

The tests for chlorine are the same as for HCl, to which the reader is referred.

CHROMATES AND CHROMIC ACID.

(CrO_3.)

When a solution of bichromate of potash is decomposed by concentrated sulphuric acid, a red deposit, consisting of beautiful crimson-red needles is produced, which is chromic acid (CrO_3). These red needles are deliquescent, and produce a brown viscid liquid. Chromic acid has no odour, it has an acid taste; it is soluble in water, the solution being brown, and the colour is very diffusible; the highly diluted solution is brown with a tinge of green, and is slowly decomposed under the influence of solar light, oxygen being disengaged, and chromate of sesquioxide of chromium deposited. Chromic acid tinges the skin yellow, and the stain is removed by alkalies. Ignition decomposes it into oxygen, and sesquioxide of chromium.

The chromates, with a weak base, are decomposed by ignition, the chromic acid yielding half of its oxygen, and being thus transformed into sesquioxide of chromium.

All the chromates of alkalies, and those of lime and of magnesia, of nickel, of zinc, &c., are soluble in water, and if an acid be added, and then alcohol, sugar, and any other organic substance, the chromic acid (CrO_3) of the chromate is converted into sesquioxide of chromium (Cr_2O_3), so that the liquor becomes of a green colour.

But even without the addition of any organic matter, all the soluble chromates are decomposed by boiling with HCl, and chlorine is disengaged; thus—

$$2KO,CrO_3+8HCl=$$
$$2KCl+Cr_2Cl_3+8HO+3Cl.$$

Sulphurous and hydrosulphuric acid decompose chromic acid also. The first producing sulphate of sesquioxide of chromium; thus—

$$2CrO_3+3SO_2=$$
$$Cr_2O_3,3SO_3.$$

The second produces water and sesquioxide of chromium, mixed with sulphur; thus—

$$2CrO_3\ +3HS=$$
$$3Cr_2O_3+3HO+3S.$$

For the mode of analysing the insoluble compounds of chromium, *see* CHROMIUM (Salts of Chromium).

The neutral chromates of alkalies and their aqueous solutions are yellow. The acid chromates, or bichromates, are orange-red. The colour of these solutions is highly diffusible. The yellow, or greenish-yellow colour of the neutral chromates is changed into orange-red when a mineral acid is added thereto.

Heated with sulphuric acid, the chromates and chromic acid are decomposed, oxygen being disengaged, and sulphate of sesquioxide of chromium formed.

TESTS AND REACTIONS.

HS. . . . . In perfectly neutral solutions, a
Greyish blue. precipitate is produced; *but if the solution* contains a free acid, the liquor becomes of a
Green colour. and a deposit of sulphur takes place. The greyish-blue precipitate consists of a mixture of sulphur, and of oxyde of chromium. When a free acid is present, the precipitate consists of nothing but sulphur; in either case the liquor is green, and, con-

		tains sulphuric acid formed by the oxydisation of the sulphur of the HS at the expense of the oxygen of the chromic acid. (See Table IV., Observation *l*; Table XV., Observation *e*; Table XXIII., Observation *h*.)
SO_2.		Reduces chromic acid into sesquioxide of chromium which remains dissolved in the liquor, and the latter assumes a
	Green colour .	Sulphuric and hyposulphuric acids are formed at the same time.
$\bar{O},HO$ . . $\bar{T},2HO$ . . $\bar{Ci},3HO$ . . HCl + alcohol.	. . . .	All these organic acids, and hydrochloric acid especially, with addition of alcohol, and the help of heat, reduce chromic acid, and the chromates, to the state of sesquioxide of chromium (Cr_2O_3), the liquor becoming of a
	Green colour .	Muriatic ether, mixed with aldehyde, is disengaged pending the boiling.
BaCl. . . .	*Pale yellow* .	precipitate (BaO,CrO_3), soluble in HCl and in NO_5.
$PbO,\bar{A}$ . .	*Yellow or orange*	. precipitate (PbO,CrO_3): soluble in caustic potash, and sparingly in NO_5. The addition of NH_3, with the help of heat, converts this precipitate into basic chromate of lead, of an *orange-red colour*.
AgO,NO_5. .	*Dark red* .	precipitate, which boiling renders purple-black (AgO,CrO_3), the precipitate is soluble in NO_5, and in NH_3.
NaO,CO_2 + KO,NO_5.	. . . . .	Fused with carbonate of soda and nitre, all the insoluble chromates are decomposed, and the fused mass, on being treated with water, produces a
	Yellow . .	solution. If a few drops of NO_5, or other mineral acid, are added, the solution becomes orange red. (See Table I—B., Observation *p*.)

SO_3,HO.		Heated in a test-tube with concentrated SO_3HO, oxygen is evolved, and, consequently, a red-hot splinter is
	Rekindled . .	when plunged into the tube. (See Table I.—C, Observation *j*.)

Blowpipe.—Chromates fused with borax before the blowpipe, yield a bead of an *emerald green* colour.

In the tests for bases it will be seen that ammonia precipitates completely oxyde of chromium from its solutions, with the help of heat; but when the chromic acid of chromates has been reduced to the state of sesquioxide of chromium by non-volatile organic substances, such as tartaric or citric acids, &c. NH_3 produces no precipitate.

The best tests of chromic acid are HS and SO_2, fusion with NaO,CO_2+KO,NO_5—$PbO,\dot{\bar{A}}$ and the emerald green bead produced by fusion with borax before the blowpipe.

CHROMIUM.

Metallic Chromium is in hard greyish-white brittle masses, requiring the highest heat of a blast-furnace for fusion. It is hard enough to scratch glass, and is capable of receiving a fine polish. Its specific gravity is 5·9. When pure, it is not magnetic, nor does it oxydise, at least at the ordinary temperature, but at 5°, either above or below 0° Fahr. it is said to have a sensible action on the magnetic needle, and at a low red heat it absorbs oxygen, and becomes converted into sesquioxyde (Cr_2O_3). It is scarcely attacked by the most powerful concentrated acids, yet hydrofluoric acid can dissolve it, hydrogen gas being evolved at the same time; the alkalies, on the contrary, easily oxydise it, especially when mixed with chlorates or nitrates, the result of the fusion being an alkaline chromate.

SESQUIOXYDE OF CHROMIUM AND ITS SALTS.

There are several oxides of chromium, but the only important one is the sesquioxyde.

Sesquioxyde of chromium, after ignition, is of a fine grass-green colour. Its hydrate is of a greyish-green colour, and is readily soluble in acids, but after ignition they have scarcely any action upon it. It is, however, only after having been strongly heated that it thus becomes insoluble, for if it has been only dried at a heat short of ignition, it dissolves in acids, though with difficulty.

The insoluble compounds of chromium may be analysed in the following manner:

Reduce the compound in as fine a powder as possible, and mix it with three or four times its weight of KO,NO_5, or of KO,NO_5 mixed with NaO,CO_2, and calcine the mass at a bright red or white heat in a well-covered hessian crucible. It is frequently necessary to continue the heat for a couple of hours. Reduce the crucible and its contents, after it has cooled down, into coarse powder, and digest it in boiling water. If any chromium was present, the water will assume a more or less deep yellow colour owing to the chromate of alkali which it will have dissolved. The addition of a little NO_5 to slight supersaturation produces a clear orange-yellow solution, owing to the alkaline neutral chromate being thus converted into a bichromate. Solution of $Pb,O\bar{A}$ added to that solution, produces then a *bright yellow* precipitate of chromate of lead, and on heating another portion of the yellow solution with HCl and alcohol it becomes green, owing to the conversion of the chromic acid into sesquioxyde of chromium.

The salts of sesquioxyde of chromium are green, or of a deep violet colour or red; their solutions are always of a blackish green, or of a bluish colour. All the salts of chromium, the acid of which is volatile, are decomposed by a red heat.

TESTS AND REACTIONS.

NH_4S.		In neutral solutions, a
	Greenish . .	precipitate of Cr_2O_3,HO, mixed with a little Cr_2S_3. (See Table XVIII., Observation *d*.)
NH_3. . . .	*Greyish-blue* .	precipitate somewhat soluble in the reagent: hence a dingy-green colour, with a purple hue, when viewed through transmitted light. If, however, an excess of NH_3 is added, and the whole be boiled, precipitation is complete.
KO. . .	*Bluish-green* .	precipitate (Cr_2O_3, HO) very soluble in an excess of the reagent, the precipitation is complete if some NH_4Cl is added to this potash solution. (See Table V., Observation *s*, Table XVIII., Observation *h*.)

Fused with KO,NO_5, yellow mass (KO,CrO_3), which dissolves in water, communicating to it a highly diffusible yellow colour (See Table I., B, Observation *p*).

Blowpipe.—Fused with microcosmic salt or borax upon a hook of platinum wire, either in the inner or outer flame of the blowpipe, an emerald-green bead is produced. (See Table V., Observation *t*.)

CITRATES AND CITRIC ACID.

($C_{12}H_5O_{11}$,3HO), or ($\overline{Ci}$,3HO.)

Citric acid is a tribasic acid, that is to say, it is an acid which requires three equivalents of base to produce neutral salts ; these three equivalents of base may be partially replaced by a corresponding number of equivalents of water.

Citric acid crystallises in colourless, transparent, rhombic prisms. It has a sour and agreeable taste ; it is very soluble in water, but is insoluble in ether. Its aqueous solution be-

comes mouldy by keeping. Heated to redness it becomes charred, and pungent acid vapours are at the same time disengaged, which however have not the odour of those produced by tartaric acid. It is very soluble also in alcohol, but insoluble in ether. Heated with NO_5, citric acid is transformed into oxalic acid.

The alkaline citrates, and those of weak bases are soluble in water, and their solution, like that of citric acid, prevents the precipitation of alumina, peroxyde of iron, and of protoxyde of manganese, by alkalies. The citrates of the alkaline earths and of most metallic oxydes are insoluble, or only sparingly soluble in water. (*See* Table XXII.—A, Observation *e, h.*)

TESTS AND REACTIONS.

$CaCl$, ...	. . .	in solutions of *free citric acid* $CaCl$ produces
	Nothing	but in those of citrates, a
	White . .	precipitate $(3CaO)\overline{Ci}$, basic citrate of lime, is formed, which is insoluble in KO, but immediately soluble in sal-ammoniac, wherefore, if the solution contains ammonia, or if the free citric acid be neutralised with ammonia, the addition of $CaCl$ will produce no precipitate *in the cold;* but, on boiling, a white basic citrate of lime is precipitated. (See Table XXII —A. Observation *j*.)
$CaO, + Aq$		(Lime-water) even in excess, produces
	No precipitate .	in *the cold,* in solutions of citric acid, and of citrates, but on boiling with an excess of the reagent, basic citrate of lime $(3CaO)\overline{Ci}$, in the form of a
	White. . .	powder, is deposited, which almost totally disappears as the liquor cools.
$PbO, \overline{Ci}$ . .	*White* . .	precipitate; sparingly soluble in ammonia, and in salts of ammonia, but immediately soluble in a solution of citrate of ammonia.

All these tests are quite characteristic.

COBALT.

Cobalt is a brittle metal, of a white colour, with a slight reddish hue, very difficult to obtain in a state of purity, as it is almost invariably found to contain traces of iron, of nickel, of arsenic, and of carbon. It is hard and capable of receiving a fine polish; its fracture presents very fine grains, like that of cast steel; its specific gravity is 8·5 or 8·6. It is difficult to fuse, its melting point being nearly as high as that of iron, and it is slightly magnetic; it is not altered by exposure or by contact with water at the ordinary temperature, but at a red heat it oxydises pretty rapidly. Sulphuric and hydrochloric acids act upon it, though slowly, and the action is promoted by heat, hydrogen gas being evolved; with concentrated sulphuric acid, sulphurous acid is evolved; nitric acid however dissolves it rapidly, and so does aqua-regia. The concentrated solutions of cobalt are blue when pure, or green when impure, either of these solutions becomes pink by dilution with water, though a small amount of impurity, especially of iron or of nickel, greatly impairs the purity of the pink colour, and imparts to it a dingy brownish hue proportionate to the amount of these metals which may be present.

OXYDE AND SALTS OF COBALT.

(CoO.)

Pure anhydrous protoxyde of cobalt is a powder of a dark and dingy olive-green colour. The beautiful ultramarine blue precipitate produced by pouring a slight excess of KO in solutions of cobalt, is not a hydrated oxyde, but a basic salt of cobalt, which becomes of a flesh colour, greenish, and finally of a dingy greenish-brown colour, which is an intermediate oxyde (Co_3O_4,HO).

The protosalts of cobalt, in the anhydrous state, are blue;

in the hydrated state, or in concentrated solution, in dilute solution they have a beautiful pink colour, which is characteristic. If the solution is concentrated, it is blue, as we just said, but the slightest quantity of iron renders such concentrated solutions green. All the salts of cobalt have an acid reaction, and are decomposed by a red heat.

TESTS AND REACTIONS

HS.	*No precipitate*	in acid solutions; but the neutral solutions, when the acid of the salt is an extremely weak one, are imperfectly precipitated
	Black . . .	by this reagent. If, however, the solution contain an excess of $KO,\bar{A}$, or other alkaline acetate, the whole of the cobalt may be precipitated by HS.
NH_4S.		In neutral and in alkaline solutions,
	Black . . .	precipitate (CoS), insoluble in excess and in alkalies; difficultly soluble in HCl.
KO. . . .	*Ultramarine blue* .	Gelatinous precipitate, of a beautiful colour, which becomes
	Greenish . .	by exposure, and reddish by boiling. Organic matter interferes with the production of this precipitate. (See Table V., Observation *m.*)
NH_3		A slight quantity of NH_3 produces a
	Blue . . .	precipitate like KO; but an excess of NH_3 redissolves it into a *dingy green liquid* which becomes *brown* by exposure. If ammoniacal salts are present in sufficient quantity no precipitate at all is produced. (See Table XXXI., Observation *a.*)
KCy. . . .	*Brownish* .	precipitate (CoCy); in acid solutions, soluble in an excess of the reagent, with the help of heat, which solution is not reprecipitated by an acid.

Blowpipe.—With borax on the hook of a platinum wire,

both in the inner and outer flame, a bead of a splendid blue colour is produced. If, however, too great a quantity of the substance has been fused with the flux, the bead may have such a deep blue colour that it appears quite black; but its blue colour may always be rendered evident by smashing the bead while it is red-hot and soft, and then looking at it when cold by holding it against the light; or else, the fused bead may be shaken off, and the small quantity which remains adhering to the platinum wire being fused again with a fresh quantity of borax, a bead of a more or less intense blue colour will be obtained. This test is the most characteristic. Remember, however, that copper, or the salts of copper, produce with borax in the oxydising flame of the blowpipe, a bead which is green when hot, but which becomes *pale blue* on cooling. In the reducing flame, however, the bead of copper is brownish-red. (See Table XXIV., Observation *g*.)

COPPER.

(Cu.)

Copper has a well-known brownish-red colour, and is capable of receiving a very high polish. Rubbed between the fingers it has a characteristic odour known to everybody; it is very malleable and ductile, harder than either silver or gold, excellent conductor of heat and of electricity, and it has more tenacity than any other metal except iron. Its specific gravity is from 8·66 to 8·72; copper-wire has a density of 8·96; copper melts at a bright red heat, or at about the same temperature as gold, and whilst in a state of fusion it is slightly volatilised so that a fine green colour is imparted to the flame. It is not altered at all at the ordinary temperature by exposure to either atmospheric air or even oxygen, provided these gases be dry, but in a moist atmosphere it becomes coated with a green crust, which is principally a

hydrated carbonate of copper. Heated to redness in contact with the air, it becomes covered with black scales, which are oxyde of copper (CuO), and which easily peel off on cooling.

Dilute sulphuric acid and *Hydrochloric acid* have scarcely any action on metallic copper, even with the help of heat; but boiling oil of vitriol dissolves it, with evolution of sulphurous acid.

Nitric acid attacks metallic copper with great energy, producing nitrate of copper, and there is a disengagement of nitric oxide, which, in contact with the air, is changed into suffocating ruddy fumes.

Aqua regia dissolves metallic copper rapidly.

Organic acids, oils, and *greasy* substances oxydise metallic copper pretty rapidly.

Ammonia with the contact of the air dissolves copper, and the solution acquires then a magnificent blue colour.

Dilute, but not concentrated, solutions of *common salt* dissolve copper rapidly.

A very small proportion of either *arsenic* or *phosphorus* is sufficient to render copper white and brittle.

OXYDE AND SALTS OF COPPER.

(CuO.)

Oxygen combines with copper in various proportions, but the only combination of interest here is the protoxyde, CuO.

Dry oxyde of copper (CuO) is a black powder. Its hydrate (CuO,HO) is *light blue.* The solutions of its salts have a *blue colour,* but the salts themselves, in the solid or crystallised state, are blue, emerald green, or, in the anhydrous state, they are *white* or *brown.* The solutions of copper are always more greenish when they contain an excess of acid. The neutral insoluble salts are blue.

TESTS AND REACTIONS.

HS . . .	*Black, brownish-black* . .	precipitate (CuS). Should the liquor contain a great excess of acid, it must be diluted with water, or the great excess of acid should be neutralised with NH_3, for otherwise HS will produce no precipitate. (See Table XXIV., Observation *h*.)
NH_4S . . .	*Black* . .	precipitate (CuS), slightly soluble in an excess, and entirely soluble in KCy. (See Table XXX., Observation *b*.)
KO . .	*Bulky light blue* . .	precipitate (CuO, HO), which, in concentrated solutions, or by standing or boiling, turns *black*, owing to its becoming dehydrated. Certain organic substances, and especially $\bar{T}$,2HO, interfere with the formation of this precipitate; but the liquid assumes a blue colour, as if NH_3 had been used. (See Table XVII., Observation *d*.)
NH_3	*Greenish* . .	A very small quantity produces a precipitate (Basic salt), immediately soluble in a small excess of the reagent, forming a splendid blue liquor, which is a solution of a basic double salt of ammonia, and oxyde of copper.
KO,CO_2 . .	*Blue* . .	precipitate (CuO,CO_2), which becomes *black* by boiling.
NH_4O,CO_2 .	*Greenish* .	precipitate, if only a small quantity of reagent has been employed; but which immediately dissolves, and becomes intensely
	Blue . . .	if an excess of it be added.
KO,CrO_3 . .	*Brownish-red*	precipitate (CuO, CrO_3).
K_2,Cfy . .	*Crimson* or *brown* . .	precipitate (2CyFe,)4CyCu); insoluble in HCl. (See Table IV., Observation *r*.)

Metallic iron.—In concentrated solutions, scales of metallic copper fall down; in dilute solutions, the bar of iron becomes quite coated with a film of copper. If the quantity of copper thus deposited upon the iron, is too small to be visible, the bar should be wetted with a solution of sal ammoniac and then exposed to the flame of a spirit-lamp, which then assumes a beautiful and characteristic green colour. This green colour is well seen only in the dark, or in a somewhat dark place, such as, for example, by holding the lamp in the shadow of a table.

Blowpipe.—Mixed with NaO,CO_2, or better still, with KCy, and heated in the inner flame of the blowpipe upon a charcoal support, the compounds of copper are reduced; and on removing the slag, or the portion of the charcoal on which it has been fused, and triturating it with water in an agate mortar, and elutriating, spangles of metallic copper are left behind, which in case of doubt may be at once identified by pouring a few drops of NO_5 upon them, and testing the solution thus obtained with NH_3, which will produce the characteristic blue colour, and after supersaturating the NH_3 with an excess of acetic acid, testing the solution with K_2Cfy, will produce the characteristic crimson precipitate. (See Table I.,—B, Observations *i*, *n*.)

CYANIDES. (See HYDROCYANIC ACID.)

CYANOGEN.

(C_2N or Cy.)

Cyanogen is a colourless gas, which at the ordinary temperature may be converted into a thin liquid by a pressure of about four atmospheres. It has a pungent characteristic odour, resembling in some measure that of peach blossoms or bitter almonds; its specific gravity is 1·806. Water dissolves five or six times its volume of this gas, wherefore it

must be collected over mercury. Alcohol may dissolve twenty-five times its bulk of it. Both the aqueous and alcoholic solutions, however, suffer spontaneous decomposition when left exposed to daylight, a black substance being produced, which may be represented as composed of cyanogen and water, and the solution is found to contain carbonate of ammonia, cyanide of ammonium, oxalic acid, and urea. Cyanogen gas is inflammable, and burns with a pink flame; this colour of the flame and the odour of the gas are quite characteristic properties, as they are possessed by no other gas. A mixture of cyanogen and of oxygen, being inflamed, detonates. Cyanogen is absorbed by solution of potash, the result being the production of an alkaline cyanide mixed with some cyanate of the alkali. The presence of cyanogen in such a solution may, therefore, be easily detected by the reagents employed for detecting cyanides.

DITHIONIC ACID. (S_2O_5.) (See HYPOSULPHURIC ACID.)

FERRICYANIDES.

(M_3*Cy_6Fe_2)

Ferricyanides may be considered as the result of the combination of sesquicyanide of iron (Fe_2Cy_3) with other cyanides, and from which a peculiar acid, called hydrocyanoferric acid, may be obtained. That acid may be considered as a combination of 3 equivalents of hydrogen with another hypothetic radical called ferricyanogen (Cy_6Fe_3) which, combining with 3 equivalents of metal, produces, according to this hypothesis, ferricyanides ($M_3Cy_6Fe_2$).

Ferricyanides accordingly may be considered as consisting of metallic cyanides combined with sesquicyanide of iron, or as consisting of metals combined with ferricyanogen.

Like ferrocyanides, ferricyanides do not exhibit any of the

* M_3 stands for 3 equivalents of the metal of the ferricyanide.

properties of cyanides, or of the salts of iron, except, as is the case with ferrocyanides, by submitting them to a peculiar treatment, since, like the latter, the elements of the cyanogen which they contain, are combined with iron, so as to form a peculiar compound radical. Ferricyanides are not poisonous.

Ferricyanide of potassium produces with metallic solutions precipitates which often possess a characteristic colour, and in which the 3 equivalents of potassium are replaced by 3 equivalents of the precipitated metal. (See Ferricyanide of Potassium, in the Dictionary of Reagents.)

FERROCYANIDES.

(M_2*Cy_3Fe.)

Ferrocyanides are compounds which, treated by acids in presence of water, yield hydrocyanoferrous acid (H_2Cy_3Fe), which may be considered as resulting from the combination of 2 equivalents of hydrogen with a hypothetic radical (Cy_3Fe), ordinarily called Ferrocyanogen; and it is this hypothetic radical which, when combined with 2 equivalents of metal, forms the compounds called Ferrocyanides. Ferrocyanides may be considered as consisting of metallic cyanides combined with protocyanide of iron, or as consisting of metals combined with ferrocyanogen.

Admitting the latter hypothesis, it is not surprising that these compounds should differ in their general properties, and in their behaviour with tests, from those by which cyanides and salts of iron are usually recognised, since the elements of cyanogen which they contain, are combined with iron so as to form a peculiar complex or compound radical.

All the alkaline and earthy ferrocyanides are soluble in water, and are more or less yellow. They are neutral to test papers, have no odour, and are not poisonous. They are

* M_2 stands for the *metal* of the ferrocyanide.

not altered by boiling with water, nor by exposure to the air, and ignition destroys only the cyanide of iron, an alkaline cyanide being left. In order to detect the presence of iron in ferrocyanides, it is necessary to decompose them completely.

Ferrocyanides are decomposed when heated with concentrated SO_3,HO, carbonic oxyde being evolved and sulphate of ammonia (NH_4O,SO_3) formed. (See Table I.,—C, Observation *i*, § 3.)

The solution of metallic ferrocyanides may be detected by the precipitates which they give with the solutions of the various metallic oxydes, and especially by the dark blue precipitate which they yield when tested with the solution of a persalt of iron; for example, with Fe_2Cl_3. (See Ferrocyanide of Potassium in the list of tests.) Ferrocyanides may also be treated by hot fuming nitric acid, to which HCl is subsequently added; a solution is thus obtained, in which peroxyde of iron may be easily detected in the usual way. (See Ferrocyanide of Potassium, in the Dictionary of Reagents.)

FLUORINE.

(Fl.)

Fluorine has never been obtained, or at least not so as to be examined. It is probably a gas of a yellowish colour, like chlorine. Its properties are not known.

FLUORIDES. (See HYDROFLUORIC ACID.)

FORMIATES AND FORMIC ACID.

(C_2HO_3,HO, or $\overline{Fo}O_3,HO$.)

Formic acid is a clear, colourless, liquid which results from the action of all oxydising bodies upon organic matter. Its name is derived from its occurrence ready formed, in the

animal kingdom, namely in the body of ants. In the concentrated state formic acid emits slight fumes in the air; it has an acid, penetrating smell, and at a temperature below 32° Fahr. it crystallises in very shining scales. Like acetic acid, it is a powerful acid which blisters the skin, and produces then a very painful sore. It boils at 212°, its vapour is inflammable, and burns with a blue flame. Formic acid dissolves in all proportions, in water and in alcohol.

All formiates are soluble in water. The formiates are decomposed by exposure to a red heat, carbonic acid, carburets of hydrogen, and water being disengaged, whilst a residue of carbon and of a metallic oxyde, or of a metal, is left. If the formiate is one of alkali, the residue is an alkaline carbonate. Heated with an excess of SO_3,HO, all formiates are decomposed, pure CO being disengaged, and HO is formed, which unites with the sulphuric acid.

TESTS AND REACTIONS.

$Fe_2Cl_3 + NH_3$		behaves with formic acid and formiates exactly as with acetic acid, and acetates; that is to say, Fe_2Cl_3 added to a solution of formic acid is
	Not precipitated . .	by NH_3, if that reagent be added only in sufficient quantity to neutralise the solution, which becomes of a
	Dark red colour,	but no precipitate is produced; if however, more NH_3 be added, or if the liquor be boiled, a
	Reddish-brown . .	precipitate is produced, which is a basic salt of iron.
AgO, NO_5. .	*No precipitate* . .	in free formic acid; but in pretty concentrated solutions of formiates a
	White . . .	precipitate ($\overline{AgO,FoO_3}$) is produced, which *in the cold* becomes
	Black or *grey* .	by standing. If, however, the liquor be heated, the
	Black or *grey* .	precipitate, which consists of metallic silver, is at once produced, even in dilute solutions of free formic acid.

The reduction of the silver is attended with an evolution of CO_2, thus $C_2HO_3,HO + 2AgO,NO_5, = 2Ag + 2NO_5HO + 2CO_2$.

SO_3,HO . . . . Heated with concentrated SO_3,HO formic acid and formiates are decomposed without being blackened, carbonic oxyde being disengaged, may be kindled, and it then burns with a

Blue flame . without smoke; no *carbonic acid* is given off, as is the case when oxalic acid or oxalates are treated in the same manner. (See Table I.,—C, Observation *i*, § 2; and Table XXVII.,—B, Observation *b*.)

HgCl . . . . Heated with a solution of corrosive sublimate (HgCl), the latter is reduced to the state of Calomel (Hg_2Cl), hydrochloric and carbonic acids being disengaged at the same time.

Hg_2O, NO . *No precipitate* with *free* formic acid.

White . . . precipitate in concentrated solutions of the formiates of alkali, this precipitate is sparingly soluble, and becomes *grey* after a short time, because a reduction of metallic mercury takes place. By boiling, this reduction is effected, even in dilute solutions.

The most characteristic test for formic acid is the reduction of the metal in the solutions of the salts of mercury and of silver.

GALLATES AND GALLIC ACID.

($CyHO_3$, 3HO, or $\bar{G}$, 3HO.)

Gallic acid may be obtained in voluminous crystals; but that which deposits from a boiling solution is in fine, silky

needles of a bright buff colour. Gallic acid is sparingly soluble in cold water, 100 parts of which dissolve only 1 part of gallic acid; but it is abundantly soluble in boiling water, 3 parts of which dissolve 1 part of the acid. Its solution, at first, becomes yellow by exposure; then it turns green, purple, and finally black. It has a sour, astringent taste. It is very soluble in alcohol, but sparingly so in ether.

GLUCINIUM.

Glucinium is a metal which is said to resemble aluminium very much. It is not oxydised by exposure. It has no action on water at a boiling temperature; but if the metal be heated to redness in contact with the air, it burns very vividly. Concentrated SO_3,HO dissolves glucinium, with disengagement of SO_2; but the dilute acids, and solution of caustic potash dissolve it, and while it is dissolving a disengagement of hydrogen takes place. Nitric acid dissolves it also, and nitric oxyde is evolved.

GLUCINA AND SALTS OF GLUCINA.

(Gl_2O_3)

Glucina, or Sesquioxyde of Glucinium, possesses many of the characters of alumina, being, like the latter oxyde, of a white colour, infusible and insoluble in water. Its specific gravity is 2·9, and it dissolves in the fixed alkalies, like alumina. By exposure to the air, it becomes converted into a CARBONATE, *which is not the case with alumina.* The Salts of Glucina are precipitated by potash, soda, and the alkaline carbonates.

The characteristic test for Glucina is NH_3 which produces in its solutions a white gelatinous precipitate, SOLUBLE IN AN EXCESS OF NH_4O, CO_2 whilst the white gelatinous precipitate of *alumina*, produced under the same circumstances, is *insoluble* in that reagent. (See Table XXXI., Observations *c.*)

TESTS AND REACTIONS.

CaCl . .	*Black* . .	precipitate ($CaO,\overline{\overline{G}}$), and the supernatant liquor is *dingy green.*
BaCl . . .	*Black* . .	precipitate.
Gelatine . .	*No precipitate* is produced in solutions of gallic acid, and of gallates.	
NH_3 . . .	*Brown colour* becoming *Green* . . . and sometimes *Purple.*	
Fe_2Cl_3 . .	*Bluish-black* precipitate (ink).	

The reaction with *gelatine* is sufficient to distinguish gallic from tannic acid, the latter producing a precipitate with gelatine; so that if a fragment of skin be immersed in a solution of tannic and of gallic acids, the first of these acids will be absorbed, whilst the gallic acid will remain in solution. If, however, gallic acid be kept in a state of fusion at a temperature of about 400° Fahr. it undergoes a modification, and acquires the property of precipitating gelatine.

GOLD.

(Au.)

Gold is a well-known metal, of a yellow colour when in masses; in fine powder it has a violet tinge, but friction restores its yellowish colour. It is tasteless, odourless, and more malleable than any other metal. Its specific gravity is (cast) 19·258, (hammered) 19·367. It melts at a temperature a little above the fusing point of silver, and it contracts more than any other metal in passing from the liquid to the solid state. It is not affected by air, oxygen, or water, at least at the ordinary temperature ; no acid or alkali, singly, has any action upon it; we must, however, except selenic acid which attacks it and becomes thus converted into selenious acid. Nitric acid *alone* has no action upon gold; but when mixed with HCl—HI—HBr &c., an *aqua regia* is produced, which

dissolves gold, and transforms it into chloride, iodide, or bromide of gold. In fact, all the acids which, like chromic acid and selenic acid, can disengage chlorine when mixed with hydrochloric acid, dissolve gold. Chlorine attacks gold even in the cold, and gold leaf dissolves rapidly in an aqueous solution of chlorine. Bromine has the same property, but iodine has no action upon it.

PEROXYDE AND SALTS OF GOLD.

(AuO_3.)

Peroxyde of Gold in the dry state is a brown powder, the hydrate which has been left to dry spontaneously has a chesnut-brown colour ; the wet hydrate is orange-yellow. Peroxyde of gold is reduced to the metallic state by exposure to even diffuse daylight. Heat evolves oxygen from it, and metallic gold is left.

The salts of gold are yellow or reddish, and always have an acid reaction. All the salts of gold are decomposed by a red heat, metallic gold being left. Almost all the organic acids reduce gold, but none so readily as oxalic and formic acids.

TESTS AND REACTIONS.

HS . . .	*Black* . .	precipitate (AuS_3), insoluble in the simple acids, soluble in aqua regia, and in a great excess of NH_4S. This black sulphuret of gold is produced by HS, even in very acid solutions.
NH_4S . . .	*Black* . .	precipitate (AuS_3), soluble in an excess of the reagent, and in aqua regia. Insoluble in the simple acids.
NH_3 . . .	*Yellow*, or *reddish-yellow* .	precipitate, in moderately concentrated solutions (fulminating gold). If the liquor is acid, or contains ammoniacal salts, heat must be applied; for, in

		such cases, *no precipitate* is produced in the cold. Certain organic substances interfere with the production of this precipitate.
KO . . .	*Reddish-yellow* .	precipitate, in concentrated and *neutral* solutions, especially with the help of heat,
	Nothing	in acid solutions.
K_2,Cfy		Fine emerald green colour.
$\bar{O}$,HO		Brown precipitate of metallic gold.
FeO,SO_3		In very dilute solutions a
	Bluish colour,	is produced, but after a while a
	Brown . . .	precipitate of metallic gold falls down, which, when rubbed in an agate mortar, assumes the metallic lustre of gold. If the solution is not too dilute, the brown precipitate is immediately produced. (See Table IV., Observation *u*. Table XVI., Observation *h*.)
SnCl		which *contains a small quantity of* $SnCl_2$, or to which enough HCl has been added to render it clear, produces a fine
	Red purple .	precipitate (purple of cassius) If the solution is very dilute, the purple precipitate has a less red tinge. The hue of that precipitate varies much.
Metallic zinc	*Brown* . .	precipitate (voluminous), which is metallic gold.
Hg_2NO_5 . .	*Black* . .	precipitate.
Infusion of Galls. .	*Black* . .	precipitate (metallic gold), which, with the help of heat, becomes brownish yellow.

The most characteristic tests for gold are FeO,SO_3—$\bar{O}$HO—HS—SnCl.

HYDRIODIC ACID.—IODIDES.

(HI.)

Hydriodic acid is a colourless gas, which emits thick fumes when exposed to the air; it has a pungent, irritating odour, and is as soluble in water as hydrochloric and hydrobromic acids, which it resembles in many respects. Its aqueous solution undergoes spontaneous decomposition by contact with the air, and becomes of a brownish red colour, which is due to a separation of iodine which dissolves in the rest of the acid liquor, and gives it the brownish-red colour just alluded to. This colour however, if the bottle be well closed, disappears almost entirely, which is due to the subsequent action of oxygen upon hydriodic acid.

The Iodides of alkalies are not decomposed by being heated in contact with the air, but under the influence of a strong heat they are volatilised. Most other iodides are decomposed when heated in the air, and the decomposition is ordinarily attended with a disengagement of violet fumes of iodine. The iodides of alkalies and of earths are soluble in water, as are also those of iron, of lime, and of cobalt. Most other iodides are insoluble.

TESTS AND REACTIONS.

AgO, NO .	*Yellowish-white* .	precipitate (AgI), insoluble in NO_5, and almost insoluble also in NH_3, by which it is distinguished from chloride of silver, which is immediately soluble in an excess of NH_3. This white precipitate blackens by exposure to light.
NO_5		Decomposes hydriodic acid and iodides, their solutions acquiring a
	Yellowish-brown .	colour, accompanied by a deposit of iodine in the form of a
	Black . . .	precipitate, if the solutions are con-

centrated. If the iodide is in the solid state and heated with NO_5,

Purple fumes . of iodine will be evolved mixed with nitric oxyde. The iodine condenses on the cold sides of the tube, and may be immediately identified by the beautiful violet fumes which make their appearance on applying heat to that part of the tube where the sublimate has taken place.

NO_5 + Starch . . . . If a few drops of strong NO_5 be first poured into the solution under examination, and a *cold* solution of starch in water be further added, a beautiful

Blue colour . (iodide of starch) will be produced, which, if the quantity of iodine be at all large, will be so intense as to appear

Black, . . dilution with cold water, however, at once exhibits the blue colour of the iodide of starch. *The solution should be cold*, otherwise *no precipitate* is produced, or it speedily disappears

If the quantity of iodine present be small, the solution should be mixed with nitric acid in a flask, and a few filaments of darning cotton smeared with a little solution of starch being suspended from the stopper, or from the cork, a deep blue stain will be perceived on the cotton, after the lapse of a few hours. Heat prevents the production of the blue colour, or destroys it when produced, as we just said, but it generally reappears on cooling, though in a less intense degree. (See Table VII, Observation *p'*. Table XXVI,—B, Observation *b*)

Cl + Starch . . . . . All iodides are decomposed by chlorine, which eliminates the iodine so that the smallest quantity of an iodide may be detected in a liquid by adding thereto (in the cold) some starch

		boiled with water, and then a few drops of an aqueous solution of chlorine, which at once produce a characteristic
	Blue colour.	Take care not to use an excess of chlorine, because it would react upon the iodine and the water, producing hydrochloric and iodic acids, which have no action on starch.
$HgCl$ . . .	*Yellow* or *vermilion*	precipitate (HgI). This precipitate is yellow at first, but it very soon becomes of a magnificent vermilion colour, immediately soluble in an excess of the reagent, or of an iodide of alkali, or of HCl.
$Hg\,O,NO_5$.	*Greenish-yellow* .	precipitate (Hg_2I).
PdO,NO_5 .	*Black* .	precipitate (PdI).
SO_3,HO . .	. . .	In the solid state iodides are decomposed by concentrated SO_3,HO, especially with the help of heat,
	Violet fumes .	being evolved, sulphurous acid is formed at the same time, but if a little peroxyde of maganese, or red lead be added, gaseous iodine alone is evolved, recognisable by the violet colour of the fumes; or if in too small quantity, by the blue colour produced by suspending over it a strip of paper moistened with starch.
$KO,2SO_3$.	. . . .	Iodides in the solid state, heated in a test-tube with $KO,2SO_3$, are decomposed with evolution of
	Violet fumes.	(See Table XXVII.,—A, Observation *b*)

Blowpipe.—Heated with a bead of microcosmic salt on a copper or brass wire, in the smallest possible flame of a spirit lamp, a beautiful *emerald green colour* is produced.

The most characteristic tests for hydriodic acid and iodides are—The *blue colour* produced by solution of starch

and nitric acid; the *violet fumes* produced by concentrated sulphuric acid and peroxyde of manganese or red lead, and by heating with $KO,2SO_3$. (See also the test in Table XXVII.,—B, Observation *c*.)

HYDROBROMIC ACID.—BROMIDES.

(HBr.)

Hydrobromic acid is a colourless gas, which has much analogy with hydrochloric acid and is as soluble in water as the latter. Hydrobromic acid is immediately decomposed by chlorine, which deprives it of its hydrogen and sets bromine at liberty, the latter substance being immediately recognised by its brownish-red or hyacinth-red colour. Chlorine behaves in the same manner with the solutions of Bromides. The solution of hydrobromic acid in water undergoes spontaneous decomposition, owing to which it acquires a brownish colour due to the presence of free bromine.

Bromides have a great analogy with chlorides.

TESTS AND REACTIONS.

AgO,NO_5 . .	*White*, or *yellowish-white* .	precipitate (AgBr), difficulty soluble in NH_3, and insoluble in dilute NO_5. (See Table VII., col. 22, 23. Table XXI., Observation *v*.)
MnO_2, or PbO_2+SO_3,HO		A mixture of a bromide with MnO_2, or PbO_2 heated with concentrated SO_3,HO, evolves
	Brownish-red fumes	which are better seen by allowing them to pass through a long, narrow tube.
Cl + Ether.		Chlorine gas, or an aqueous solution of chlorine, added to a solution of a bromide, or passed through a liquor containing an insoluble bromide, for example, bromide of silver, liberates the bromine, which may be separated by means of ether.

The modus operandi is as follows:

To the liquor containing the bromide, add as much æther as will form a layer about $\frac{1}{4}$ of an inch thick, shake it well, and then add some aqueous solution of chlorine, or pass a few bubbles of chlorine gas through it, cork the whole well, shake it violently and leave it at rest for a few minutes.

TESTS AND REACTIONS.

A layer of a *Brownish red, hyacinth red,* or only of a *slight yellow colour* (according to the quantity of bromine present), will be observed floating at the top; it is an ethereal solution of bromine. If the liquor so treated is one containing the white precipitate produced by AgO, NO_5, the chloride of silver, resulting from the treatment with chlorine as just said, will sometimes float at the top of the liquor amidst the layer of ether; but after standing for a few moments, a slight shake of the tube is sufficient to precipitate it all to the bottom of the tube, and the layer of ether coloured by the bromine may be then distinctly seen.

Perhaps the most convenient way of separating the ethereal solution of bromine from the rest of the liquor is the following: The whole is to be poured into a straight pipette held vertically whilst the point is pressed either on the tip of the small finger, or on a piece of india rubber, so as to close the small aperture; after a few moments the ether holding the bromine in solution will all collect at the top of the liquor in the pipette, and by lifting up the pipette and removing the finger, the lower stratum may easily be allowed to flow entirely out so that the ethereal solution may now occupy the narrow part of the pipette down to the end, and be finally

permitted to fall into a glass test-tube, in which it may be mixed with a little solution of potash, which immediately decolorises it. The liquor is then evaporated to dryness by means of a spirit lamp, and submitted to the usual tests for the detection of bromine in bromides; namely, by treatment with MnO_2 or $PbO_2 + SO_3,HO$, or by heating it in a glass tube with $KO,2SO_3$, &c. (See Table VII. Observations *m'*, *n'*.)

TESTS AND REACTIONS.

$KO,2SO_3$	*Hyacinth red fumes*, and an *Odour of bromine*	When a bromide reduced to powder is mixed with $K,2SO_3$, also pulverised, and heated in a glass tube, ... are evolved. (For the distinctive characters between bromides and bromates, see Table XI., Observations *i*.)
Starch	*Yellow precipitate*	Produces in liquids containing free bromine a ... (bromide of starch), which however requires generally some time for its production. (See Table VII., Observations on column, 17.)

The most characteristic reactions for hydrobromic acid are those with *chlorine* and with AgO, NO_5.

HYDROCHLORIC ACID. CHLORIDES.

(HCl.)

At the ordinary temperature and pressure HCl is a colourless gas, emitting thick, white fumes; when exposed to the air, it has a suffocating odour, and a strongly acid taste. It is not a permanent gas, since, when submitted to a pressure of about 280 lb. to the square inch at a low temperature, it becomes liquid. Ordinary HCl, however, is a solution of the gas in water. The specific gravity of a saturated solution of the gas at 32° Fahr. is 1·2009, and it then contains 6 equiva-

lents of water. If such a solution be left exposed, it emits dense white fumes and loses a portion of its acid, whereby it is converted into an acid containing 12 equivalents of water, its density being then 1·128. The aqueous solution of HCl loses, at first, by boiling a large portion of its acid; but its boiling point soon becomes uniform, and it is transformed into another hydrated acid, containing 16 equivalents of water, its density being then 1·094.

Hydrochloric acid combines with all oxydes producing water and chlorides. The neutral chlorides *except those of mercury and silver*, are readily soluble in water, even chloride of lead dissolves in that menstruum in considerable quantity. Several chlorides volatilise without decomposition when strongly heated, such as, for example, chloride of tin, of arsenic, of antimony, &c.; others are completely fixed; and others are altogether decomposed by heat.

TESTS AND REACTIONS.

AgO,NO_5.. *White curdy* precipitate (AgCl) turning purple and, then black, by exposure to daylight insoluble in acids, IMMEDIATELY *soluble in free ammonia*, readily re-precipitated by saturating the NH_3 with an acid. (See Table VII, Observation *l'*) Chloride of silver fuses without decomposition into a tough, horny mass (horn silver), which is reduced into metallic silver by hydrogen, or by fusion with carbonate of soda, or with rosin.

$Hg_2O,NO_5 + NH_3$. *White*. . . precipitate (Hg_2Cl), which turns black when NH_3 is poured upon it. (See Table XIV. Observation *b*.)

$KO,2CrO_3$ + SO_3,HO . . . . Triturated with KO, $2CrO_3$, introduced into a small flask or test-tube, and treated therein with concentrated SO_3,HO, with the help of a gentle heat, a

Gas of a deep red colour (Chromate of perchloride of

chromium $CrCl_3+2CrO_3$) is produced; supposing chloride of sodium to be present, the re-action is as follows:

$$3KO,2CrO_3+3NaCl+6SO_3,HO= \\ 2CrCl_3,2CrO_3+3KO,SO_3+3NaO,SO_3+6HO.$$

By adding NH_3 to this bichromate of perchloride of chromium a *yellow liquid* is produced (NH_4O,CrO_3,) which becomes red by the addition of an acid ($NH_4O,2CrO_3$).

A characteristic test in the dry way is the beautiful BLUE COLOUR, which is imparted to the flame of a spirit lamp, when the wick being as low as possible, a particle of the compound under examination, previously mixed with microcosmic salt, *on a brass wire* is held in that flame. All chlorides, and, in fact, all compounds containing chlorine, exhibit this phenomenon. Bromides exhibit the same reaction, but the colour of the flame has a greenish tinge.

All these tests are quite characteristic; nitrate of silver however, is the most important, since the white precipitate which may be produced with AgO,NO_5 by other acids, is soluble in NO_5, bromides and bromates, cyanides, iodides and iodates being the only exceptions; but chloride of silver is more readily soluble in NH_3 than bromide, bromate, and cyanide of silver, whilst the iodide of that metal is altogether insoluble therein.

HYDROCYANIC ACID. CYANIDES.

(HCy.)

Pure hydrocyanic acid is a volatile, colourless, inflammable liquid, soluble in all proportions in water, alcohol and ether. Its aqueous solution in a diluted state has a bitter taste, and an odour which, to a certain extent, resembles that of bitter almonds, than which it is much less fragrant and aromatic,

and is accompanied with a peculiar acrid sensation. Pure hydrocyanic acid burns with a bluish flame, like alcohol; it undergoes spontaneous decomposition, becoming darker and darker, and eventually leaves a black, solid mass, the composition of which is as yet imperfectly known. It is one of the most violent poisons, one drop of the concentrated acid put in contact with a mucous or abraded surface, being sufficient to destroy animals of considerable size. The antidotes for this fearful poison are chlorine and ammonia, but it is rare that these substances can be administered soon enough to be of use. It is well to pour some ammonia in capsules, or on the floor of the laboratory, when dealing with that substance.

The cyanides of alkalies, or of alkaline earths, are soluble in water; their solution has a strongly alkaline reaction, and undergoes gradual decomposition, their odour resembles that of HCy. In the dry state they may be strongly heated without undergoing decomposition; their aqueous solution is decomposed by a protracted ebullition into NH_3, which is evolved, and a formiate which remains as residue, thus:

$$KCy \text{ or } KC_2N + 4HO = NH_3 + \underbrace{KO,C_2HO_3}_{\text{formiate of potash.}}$$

All the other cyanides, except those of gold and mercury are insoluble in water. (See Table IV., Observation *n*.) The soluble cyanides are decomposed by the weakest acids, hydrocyanic acid being disengaged. Yet it is necessary to observe that, if the acid has been left for a long time in contact with the cyanide before heating it, or if it was in very large proportion, formic acid only, instead of hydrocyanic acid, will be produced. The action of cyanogen upon iron, manganese, cobalt, and chromium, is to form compound radicals, in which the presence of these metals cannot be detected by the usual methods, Insoluble cyanides are

immediately dissolved by alkaline cyanides, and produce numerous double salts; such as, for example, the ferro and ferricyanides. Most cyanides are charred by exposure to a red heat, and the charring is frequently accompanied by an odour of bitter almonds, and sometimes of NH_3, which is due to a decomposition of the cyanogen, and of the combination water of the compound, so as to form ammonia and a formiate, *ut suprà*.

Heated with concentrated SO_3,HO, cyanides and ferrocyanides yield carbonic oxyde gas, which may be inflamed, and which burns then with a blue flame, because the elements of hydrocyanic acid arrange themselves with the elements of water, so as to produce formic acid and ammonia, which, if the SO_3,HO is in excess, are decomposed into carbonic oxyde and water; we have already alluded to this reaction in Table I.—C. Observation *i*, § 3.

TESTS AND REACTIONS.

AgO,NO_5 . . *White curdy* precipitate (AgCy), insoluble in dilute NO_5, and almost so in water, but soluble in free NH_3 (though not so readily as chloride of silver), and likewise in cyanide of potassium (KCy). The precipitate leaves metallic silver by ignition, and when drenched with HCl, an odour of HCy becomes immediately perceptible.

This white precipitate is distinguished from all others, and identified as AgCy, by first drying and then heating it in a glass tube, cyanogen gas is then disengaged, which being inflamed burns with a peach-blossom colour. (See Table XXVII.—B. Observation *a*.)

Another excellent test for identifying cyanide of silver is that proposed by M. Lassaigne, and which consists in putting into a glass test-tube, about two inches long and at most one eighth of an inch in diameter, a small fragment of potassium not bigger than a millet seed, and above that a little of the

compound supposed to be cyanide of silver; the whole is then gently heated and brought to a red heat. When the tube has quite cooled, a little water is added, then a few drops of sulphate of protoxyde of iron, and a drop of sesquichloride of iron, and, lastly, a few drops of HCl, Prussian blue, will then be immediately produced.

TESTS AND REACTIONS.

HgO . . . . Peroxyide of mercury (HgO) is soluble in HCy, and the solution cannot be precipitated by potash. If, therefore, to a solution containing HCy, potash be added, and then finely levigated HgO, if the latter dissolves in the liquor, it is a proof of the presence of HCy; peroxyde of mercury being soluble in alkaline liquids only in presence of HCy, the solution of this oxyde in the alkaline liquor is, therefore, quite characteristic of the presence of HCy.

$FeO, SO_3 + KO + HCL.$ . . FeO, SO_3, added to an alkaline solution, or to a solution which has been rendered alkaline by KO, forms a

Blue . . precipitate of Prussian blue upon *supersaturating* the alkaline solution with HCl, taking care that a slight excess of the acid be present in the liquor. The addition of KO before or after the FeO, SO_3 is immaterial, but it is absolutely necessary that an alkali be present, and also free HCl. The FeO, SO_3, should contain a little peroxyde of iron, the presence of which is secured by using a FeO, SO_3 which has been left exposed for some time to the air. The blue precipitate (Prussian blue) is insoluble in water and in HCl.

The best way of applying this test consists—

1st. In pouring some solution of caustic KO, in the liquor in which the presence of HCy is suspected, and stirring the mixture.

2nd. Some solution of ordinary FeO,SO_3, and again stirring the mixture.

3rd. *An excess* of HCl.

If HCy is present, except in certain cases mentioned below, a precipitate of Prussian blue will take place.

CuO,SO_3.	Chocolate-brown or crimson-red precipitate, according to the state of dilution of the liquor.

Cyanogen, however, cannot be detected in cyanide of mercury by any of these tests. In such a case the following process has been recommended: Put a piece of iron (an iron nail, for example) in the solution to be examined, add HCl thereto (which generally produces an odour of HCy), then potash, and finally a slight excess of HCl; a precipitate of Prussian blue will take place.

The best test, however, inasmuch as it is the most delicate, and that it is applicable to all the combinations of cyanogen, is hydrosulphuret of ammonia with excess of sulphur. It is applied as follows:—

NH_4S_5	To the solution under examination add some HCl, and put it into a small capsule. Pour into another small capsule a single drop of NH_4S, in which as much sulphur as it can take up has been dissolved, and put both capsules side by side under a small glass-beaker or jar. In the course of a few minutes (say ten minutes, though a much less time will in most instances suffice) the drop of NH_4S_5, will appear covered with a film; if the capsule be then

		withdrawn and tested with a drop of
Fe_2Cl_3 . .	a *Blood-red* . .	colour will be produced, which, if a solution of HgCl be added, will instantly disappear. If no cyanogen had been present, no sulphocyanogen would, of course, be produced, and the addition of Fe_2Cl_3 instead of a blood-red, would have produced a black precipitate of sulphuret of iron.
		This reaction is due to the fact that the polysulphurets of ammonium are instantly deprived of the excess of sulphur beyond that which constistitutes the monosulphuret (NH_4S) by cyanide of ammonium, in consequence of which sulphocyanide of ammonia (NH_4CyS_2) is produced, which in contact with persalts of iron, produces the characteristic blood-red colour of sulphocyanide of iron.

The modus operandi just described may be modified by moistening a small capsule or watch-glass with the NH_4S_5, and inverting it over another capsule, or watch-glass which contains the liquor acidified with HCl, and leaving the two little vessels the one covering the other, for a few minutes, and adding a drop of Fe_2Cl_3, to that which was moistened with NH_4S_5.

HYDROFERROCYANIC ACID. (See FERROCYANIDES).

HYDROFERRICYANIC ACID. (See FERRICYANIDES).

HYDROFLUORIC ACID. FLUORIDES.

(HFl.)

Hydrofluoric is a corrosive, colourless, and very volatile

acid, which emits thick, suffocating fumes when exposed to the air. It is soluble in water in all proportions; and it is distinguished from all other acids by its property of dissolving silica, and consequently of corroding glass; several species of glass, however, are slightly attacked by other acids, for example, by concentrated sulphuric acid, but the action of hydrofluoric acid is much more powerful. Hydrofluoric acid decomposes metallic oxydes, which it converts into metallic fluorides and water.

The fluorides of alkalies are soluble in water, those of the earthy metals and of metals are insoluble or sparingly soluble in water; and there is no fluoride of gold. The insoluble fluorides are decomposed by fusion with NaO, CO_2.

TESTS AND REACTIONS.

$CaCl$ . .	*White gelatinous*	precipitate, ($CaFl$) almost insoluble in free acids. This precipitate, however, is so transparent, that it can hardly be seen, and does not settle well, unless NH_3 be further added, which then produces a very bulky precipitate, which may be identified as $CaFl$ by mixing it with SO_3, HO, and examining its action on glass as described below.
$BaCl$ . . .	*White* . .	precipitate, soluble in an excess of HCl.
$PbO, \bar{A}$ . .	*White* . .	precipitate, soluble in free HCl.

Concentrated SO_3, HO decomposes all fluorides *with the help of heat*, hydrofluoric acid being disengaged. The best way of applying this test, consists in coating a piece of glass with a film of bees'-wax, which is easily accomplished by first heating the glass and rubbing it with a piece of bees'-wax. When the glass has become cold, lines or figures are to be traced through the wax down to the piece of glass with a pointed brass wire, a pin for example, or a pointed piece of

ivory, or of hard wood. The substance to be examined being first reduced to powder, is put into a platinum crucible, concentrated SO_3,HO is poured upon it, and the whole is then covered with the piece of glass prepared as above said, and with the writing downwards of course. The crucible so disposed is next *gently* and *carefully heated* by means of a small spirit-lamp, but at a temperature sufficiently low not to melt the wax-coating; and further, to guard against such an accident, some water should be poured, or a proper filter wetted with water should be put upon the back of the glass, in order to keep it cold. It is absolutely necessary to apply a *gentle heat* because some fluorides (for example *Fluoride of calcium*) when thus treated *in the cold* form only a viscid mass from which no HFl is evolved. A few minutes after this treatment the coating of wax is removed from the glass, by first heating it, wiping the wax off, and then washing it if necessary with essence of turpentine. If only a trace of HFl was present the tracing cannot be seen except by breathing upon the glass after having removed the wax; otherwise the glass is seen to be etched more or less deeply (see Table I.,—C, observation *c*)

HYDROGEN.

(H.)

Hydrogen is the lightest of all known substances. It is a colourless gas, inodorous when quite pure, but which, however, has generally a somewhat disagreeable smell when prepared with zinc, and more so still with iron; the odour is due to the presence of a small quantity of a hydrocarbon, of sulphuretted or of arseniuretted hydrogen. It is inflammable, and burns with a thin flame, scarcely visible in daylight. Mixed with oxygen in a bottle and inflamed, it detonates violently. If the experiment be performed in an endiometer the diminution of volume may serve, with proper precautions,

to determine the quantity of hydrogen present in the mixture, since it forms the two-thirds of the volume of gas which has disappeared after the detonation. Its specific gravity is 0·0691 or 0·0695, and therefore 100 cubic inches weigh, at the ordinary temperature and pressure, 2·14 grains only. Water is the protoxyde of that gas. Hydrogen has no action upon test papers, and is not absorbable by KO.

HYDROSULPHOCYANIC—SULPHOCYANIDES.

($CyS,H.$)

Hydro-sulphocyanic acid is a colourless, strongly acid liquid, which, unlike hydrocyanic acid, is not poisonous. It is spontaneously decomposed by ebullition and exposure to the air into various products, one of which is of a yellow colour. This acid, by its union with metallic oxydes, forms compounds called metallic *sulphocyanides,* amongst which those of the alkaline metals, of ammonium, of barium, strontium, calcium, manganese, and iron, are very soluble in water.

TESTS AND REACTIONS.

AgO,NO_5 . .	*White precipitate*	insoluble in NO_5, soluble in strong, but not in dilute NH_3.
Fe_2,Cl_3 . .	*Blood red* .	colour, but no precipitate. The blood-red colour is immediately destroyed by various substances, such as ammonia, solution of $HgCl$—$SnCl$—HS. The decolorisation produced by solutions of corrosive sublimate ($HgCl$) is employed to distinguish the blood red colour of sulphocyanide of iron, from that produced by Fe_2Cl_3 in acetic and meconic acids, in which case it is not destroyed by $HgCl$.

HYDROSULPHURIC ACID (SULPHURETTED HYDROGEN)—SULPHURETS.

(HS.)

At the ordinary temperature, and atmospheric pressure, hydrosulphuric acid is a colourless gas of a peculiar and extremely fœtid odour (that of rotten eggs). Inflamed in contact with the air, it burns with a blue flame and leaves a deposit of sulphur. Water dissolves about two or three times its own bulk of this gas, and the solution acquires thereby the characteristic odour of the gas. Alcohol dissolves about six times its bulk of it. Sulphuretted hydrogen is decomposed by chlorine, bromine, iodine, and by fuming nitric acid or aqua-regia; the decomposition is accompanied by a deposit of sulphur, thus:—

$$HS + Cl = HCl + S$$
$$HS + Br = HBr + S$$
$$HS + I = HI + S$$

Sulphuretted hydrogen is also decomposed by sulphurous acid, water being produced and sulphur deposited, thus:—

$$2HS + SO_2 = S_3 + 2HO$$

The aqueous solution of HS undergoes spontaneous decomposition, sulphur being deposited. When, however, the solution contains no atmospheric air, or is kept out of its contact, no such decomposition takes place.

Most metallic oxydes are converted by HS into *sulphurets*, with formation of *water*. The following substances are precipitated from their acid, neutral and alkaline solutions by HS.

Oxyde of silver (AgO)	Black.
Suboxyde of mercury (Hg_2O)	
Peroxyde of mercury (HgO) . . .	
Protoxyde of lead (PbO)	
Oxyde of bismuth (BiO)	
Protoxyde of copper (CuO)	
Oxyde of palladium (PdO)	
Sesquioxyde of Rhodium (R_2O_3)	Blackish Brown
Deutoxyde of Osmium (OsO_2)	Brownish Yellow.
Oxyde of Cadmium (CdO)	Yellow.
Peroxyde of Iron (Fe_2O_3) (see Table IX., observation *k*, and Table XV, observation *c*.) .	White.

The white precipitate produced in solutions of Fe_2O_3, is sulphur, and the original peroxyde of iron remains in the solution in the state of protoxyde of iron, (FeO) thus:—

$$Fe_2O_3 + HS = 2\,FeO + HO + S.$$

The following substances are precipitated from *their acid* (but not from their alkaline) solutions, by HS. (See Table IV., observation *k*.)

Peroxyde of Gold (AuO_3)	Black.
Peroxyde of Platinum (PtO_2) . . .	
Oxydes of Molybdenum MoO, MoO_2 . . .	Blackish Brown
Tellurious Acid (TeO_2)	Brown.
Deutoxyde of Iridium (IrO_2) . . .	
Protoxyde of Tin (SnO)	
Selenious Acid (by heating the liquor) (SeO_2).	Red Orange.
Oxyde of Antimony (Sb_2O_3)	
Selenious Acid (in the cold) (SeO_2) .	Yellow.
Deutoxyde of Tin (SnO_2)	
Arsenious Acid (AsO_3)	
Arsenic Acid (AsO_5)	

The preceding substances cannot be precipitated by HS from their alkaline solutions, because their sulphurets are soluble in alkaline sulphurets.

HS produces no precipitates in solutions which are somewhat concentrated and which contain a great excess of acid; in that case the solution should be diluted with a large quantity of water, or, better still, the great excess of acid should be removed by evaporation, or by partial saturation with an alkali, when such an addition is not objectionable.

The sulphurets of alkalies, and of alkaline earths, alone are soluble in water; all other sulphurets are insoluble in that menstruum. The solutions of the above-mentioned sulphurets are decomposed by acids with evolution of HS, and the liquor becomes milky or turbid, owing to a deposit of sulphur (see Table XXI., observation *e*.)

The sulphurets of the alkalies, and of the alkaline earths, alone evolve HS, when treated by concentrated SO_3,HO. Other sulphurets, such as those of iron and of manganese, do so only when water is added, or when diluted SO_3 is employed, the sulphurets of copper, of mercury, of silver, &c., disengage a mixture of sulphurous, and of hydrosulphuric acid when heated with concentrated SO_3,HO. Vermilion (sulphuret of mercury, HgS), is the only sulphuret which is not decomposed by any of the simple acids, but *aqua regia* dissolves it readily. Vermilion, when exposed to a moderate heat, becomes purple or black, but it reassumes its colour on cooling, *provided the heat applied has not been strong enough to expel any of the sulphur*, for in that case, instead of reassuming its brilliant red colour, it would remain more or less brownish, or perhaps black.

TESTS AND REACTIONS.

HCl . .	*Odour of HS*, sometimes with deposit of S and generally with effervescence (see Table I.—C, observation *m*, Table I—E, observation *f*, *h*.)
NO_5, or aqua-regia . . .	Decomposition of the sulphuret and formation of SO_3, accompanied generally by a separation of S, which continued boiling, agglomerates into *yellow lumps*.

	ruddy fumes.	With NO_5 there is a disengagement of (see Table I.—E, observation *n, o, p, r, s*; Table II.—B, observation *i.*)
AgO,NO_5, or PbO,A . .	*Black precipitate* . .	in solution of sulphurets, or of HS. Either of these two reagents may serve to identify the gas evolved, as follows —Introduce the compound to be tested into a small flask, pour upon it an excess of dilute HCl, close the flask with a perforated cork provided with a disengagement tube plunging into a beaker containing a solution of AgO,NO_5, or of $PbO,\bar{A}$, whereupon a black precipitate will be produced if HS be present.

Blowpipe.—Heated in the inner flame of the blowpipe most metallic sulphurets fuse and evolve an odour of sulphurous acid (SO_2), but this does not always take place.

If, however, the compound is mixed with NaO,CO_2 and exposed to the inner flame of the blowpipe, the fused mass on being placed upon a piece of clean silver, and moistened thereon with dilute HCl. will produce a black stain and generally evolve an odour of HS. (See Table I.—C. observation *b.*)

Exposed to a red heat in a glass tube, sulphurets sometimes yield a sublimate of sulphur, which is due to a partial desulphuration of the metallic sulphuret. (See Table I.,A, observation *q.*) Heated in a tube open at both ends an odour of SO_2 is generally evolved.

POLYSULPHURETS, OR SULPHURETS WITH EXCESS OF SULPHUR.

TESTS AND REACTIONS.

HCl.	*fœtid odour* *white precipitate*	Disengagement of HS recognisable by its always accompanied by an abundant of sulphur. If the polysulphuret is poured in the acid, it often happens that an oily yellow liquid is produced,

which is bisulphuret of hydrogen (HS_2).
Other acids produce the same reaction.

The best tests of the presence of HS or of a sulphuret are the *odour of rotten eggs* produced by treatment with HCl, and, if this is attended with effervescence, the *black precipitate* (PbS) produced by passing the gas evolved through $PbO, \bar{A}$.

Fusion with NaO, CO_2 in the inner flame of the blowpipe, and the *black stain* and *odour of HS*, produced by placing the fused mass upon a clean piece of silver and moistening it thereon with HCl. (See Table I.—C., observation *b*.)

HYPOSULPHURIC ACID (DITHIONIC ACID).

(S_2O_5)

Free hyposulphuric acid has never been obtained in the anhydrous state, and heat decomposes its hydrate into sulphurous and sulphuric acids; but it is not altered by treatment with chlorine, nor with nitric acid. It does not undergo change by exposure.

All hyposulphates are soluble in water. Heated in a glass tube an odour of sulphurous acid is evolved, because they are thereby decomposed into a sulphate and a sulphite, the heat of a spirit-lamp is sufficient for the purpose.

TESTS AND REACTIONS.

HCl		poured in solutions of hyposulphates,
	nothing	in the cold, but, on boiling, the mixture
	odour of SO_2	and production of SO_3, so that the liquor is afterwards precipitated white by BaCl.
SO_3		same as with HCl.
NO_5	*nothing*	in the cold, but by boiling
	ruddy fumes	of nitrous acid are evolved (see Table I.,—C., observation *n o*.)

The most characteristic tests of the presence of hyposulphates are the odour of SO_2, which is evolved when heated in a test-tube, and their reaction when fused with NaO,CO_2 on a charcoal support, which reaction is the same as with the sulphates.

HYPOSULPHUROUS ACID—HYPOSULPHITES.

(S_2O_2.)

Free hyposulphurous acid, undergoing almost instantaneous decomposition, is scarcely known in that state.

All hyposulphites dissolve freely in water, with the single exception of hyposulphite of barytes, which is only sparingly soluble in that menstruum. They are all decomposed by heat, those of the alkalies leave when so treated a residue of sulphate and of polysulphuret of alkali mixed.

TESTS AND REACTIONS.

HCl, or SO_3 .	Odour of sulphurous acid and *Yellowish white* precipitate, which is due to a separation of sulphur, which, however, requires often a little time to be produced *in the cold*—heat at once determines the formation of that precipitate. The solution should not be too dilute, yet even then, if heat be applied, a milkiness appears (see Table I.,—C , observation *n, o* , Table VII., observation *c;* Table XI , observation *b*.)
NO_5	produces a *Yellowish white* precipitate as in the case witk HCl —SO_3,HO, and other acids, but, by boiling, the hyposulphurous acid becomes converted into sulphuric acid, and BaO,NO_5 will then of course produce a white precipitate (BaO,SO_3). (See Table I.,—C, observation *n, o*.)
Cl, or Hypochlorites	Convert into SO_3 all the sulphur contained in hyposulphites.

I, and persalts of iron.	. . . .	Convert hyposulphites (MO, S_2O_2)* into tetrathionates (MO, S_4O_5) thus:

$$2NaO, S_2O_2 + I = NaI + NaO, S_4O_5.$$

$$Fe_2Cl_3 + NaO, S_2O_2 = NaCl + NaO, S_4O_5 + 2\ FeCl.$$

Alcoholic solution of I.	. . .	An alcoholic solution of iodine is *immediately decolorised* by solution of hyposulphites.
AgO, NO_5. .	*White* . . *yellow, yellowish-brown, brown* . . *black* .	precipitate, immediately becoming and finally The changes of colour take place in the cold. (See Table VII., observation *h.*)

Recently precipitated chloride of silver is readily dissolved by solutions of hyposulphites, and the solutions acquire a strong taste of sugar.

The most characteristic reactions are the *smell of* SO_2, *accompanied by a deposit of sulphur* when treated by SO_3 or by HCl. (See Table VII., observation *h.*)

The precipitate produced by AgO, NO_5.

The solution of recently precipitated AgCl.

The decolorisation of an alcoholic solution of iodine.

IODATES—IODIC ACID.

(IO_5.)

Iodic acid crystallises in colourless six-sided tables. It is very soluble in water; it reddens litmus paper at first, but it subsequently decolorises it. It is easily decomposed by sulphurous acid, sulphuretted hydrogen, sulphur being deposited. Exposed to a low red heat, it undergoes decomposition with disengagement of oxygen and of iodine. Most metals are oxydised by that acid. Heated with organic substances iodic acid and iodates deflagrate.

* MO stands for oxyde of metal.

The iodates of alkalies are soluble in water, nearly all other iodates are insoluble in that menstruum. Exposed to a red heat in a test-tube they are converted into iodides, oxygen being disengaged, and if it be an acid iodate, purple vapours of iodine are likewise given off.

TESTS AND REACTIONS.

BaCl . . .	*White* .	precipitate (BaO,IO_5) soluble in NO_5 (See Table XXII.—A, observation *i.*)
AgO,NO_5 . .	*White* . .	precipitate (AgO,IO_5) soluble in NH_3, but almost insoluble in dilute NO_5

When strongly heated in a test-tube or in a small glass retort, oxygen is disengaged, which may be identified as such by plunging a *glowing match* into the tube, when it will be *rekindled.* The residue is an iodide, which may be recognised as such by the reagents employed for HI and iodides. (*See* HYDRIODIC ACID.)

Transfer a portion of the dry pulverised iodates to a test-tube, apply a strong heat as just said by means of an argand spirit lamp, and throw a small piece of paper or a splinter of wood or of charcoal into the fused mass — *there will be a* DEFLAGRATION.

Put into a small test-tube a little of the well pulverised iodate with about twice its bulk of $KO,2SO_3$ and heat the whole by means of a spirit lamp, *violet fumes* will fill the glass tube and condense on its cold sides. These fumes will turn a solution of starch blue. (See Table VII., observation p'.)

IODIDES. (*See* HYDRIODIC ACID.)

IODINE.

(I.)

Iodine is generally found in commerce in the state of oblong octahedrons, or rhombohedral plates, or in scales or masses of a metallic bluish-black colour resembling plumbago

(black-lead). Its odour partakes of that of chlorine and of bromine; its specific gravity at the ordinary temperature is 4·948; it fuses at 225° Fahr., and it boils at 347° Fahr., the vapour having a most beautiful purple colour; but even at the ordinary temperature it slowly evaporates, hence the empty space of the bottles in which it is kept has a slight violet colour. It is slightly soluble in water, 7000 parts of which dissolve only 1 grain of iodine, and the solution has a brown tinge. Alcohol however dissolves it abundantly, and thus acquires a deep brown colour Iodine is precipitated from such a solution by water, in the state of a black powder. Iodine is also soluble in bisulphuret of carbon, and the solution has a bluish colour. Iodine gives to the skin a brownish-yellow stain, which is not permanent, the stain disappearing after a short time, provided the contact has not been too long, for otherwise the epidermis is destroyed. All organic substances are acted upon by iodine, their hydrogen combining with it to form hydriodic acid. The violet fumes of its vapour, and the production of a blue colour when it is put in contact with cold mucilage of starch are the most characteristic properties of iodine; this blue colour, which is due to the production of iodide of starch, disappears by boiling, but it is reproduced as the liquid cools. (See Table VII. observation *p′*.)

Iodine with the solutions of the pure and of the carbonated alkalies produces iodides of those alkaline metals; when ammonia is poured upon iodine the result is an explosive compound (iodide of nitrogen), which when dry detonates violently by the slightest pressure, or friction.

IRIDIUM.

(Ir)

Iridium when obtained by calcining the ammonia-chloride of iridium resembles spongy platinum, and assumes a metallic lustre by friction; but when in cohesive metallic masses, it has

exactly the appearance of platinum, except that instead of being ductile and malleable, it is on the contrary brittle and easily pulverised. Its specific gravity is 15·68 or 18·68; it is completely infusible and unvolatilisable. It is *insoluble in all acids*, and except it is alloyed with a large quantity of platinum, even aqua regia has only a very slight action upon it; but by fusion with $KO,2SO_3$, or with nitre, or with the alkalies, it is converted into sesquioxyde (Ir_2O_3), which by proper treatment may be dissolved in HCl. Iridium forms alloys with most metals, and especially with osmium; there is a natural alloy of iridium and osmium, which is exceedingly hard. There are four oxydes of iridium, namely IrO,—Ir_2O_3,—IrO_2,—IrO_3. The deutoxyde (IrO_2) is that which is best known. The solutions of iridium often assume a blue colour, which is supposed to be due to the presence of an intermediary oxyde of that metal.

TESTS AND REACTIONS FOR IrO_2, AND ITS SALTS.

KO or NH_3 . .	. . . . *decolorises*.	an excess of solution of KO or NH_3 the liquor, and produces a small
	black . .	precipitate but if the solution be then left exposed to the air it soon acquires a
	fine blue .	colour.
KO,CO_2 . .	*Reddish brown*	precipitate, which slowly redissolves in the liquor, which becomes gradually
	blue . . .	by exposure.
NH_4O,CO_2 .	*Blue* . .	colour by exposure.
K_2 Cfy . . .	*Decolorises*.	the liquor.
FeO,SO_3 .	*Decolorises*.	the liquor.
HS	*Decolorises*.	the liquor at first, then, after a time, a
	brown . .	precipitate is produced.
NH_4S . . .	*Brown* . .	precipitate completely soluble in excess of the reagent.
Bar of Zinc .	*Black* . .	powder of metallic iridium.
Ammoniacal salts	*Very dark brown*	powder soluble in solution of SO_2.

IRON.

(Fe.)

Pure iron is perfectly white, but it is seldom met with in that state, the iron of commerce containing always traces of carbon, of silicium, and sometimes of phosphorus and arsenic; it has a bluish-grey colour, with great metallic lustre; it is malleable, ductile, and it is the most tenacious of all metals; it has a feeble odour and taste, and becomes brittle by hammering. When iron contains a certain quantity of sulphur and of arsenic, it is more or less liable to break whilst being forged, it is then called *hot short iron* ; on the contrary, if it contain a certain quantity of phosphorus, it will break when hammered or bent in the cold, and on that account it is called *cold short iron.* There are other kinds of iron which break both whilst hot and while cold, so that they are almost useless. When iron, after having been strongly heated, is allowed to cool slowly and without being disturbed, the interior of its mass is found full of cubic or octahedric crystals, and the same effect is produced even in the cold, by submitting it to repeated and protracted short vibrations: this state greatly impairs its tenacity. The specific gravity of iron is from 7·7 to 7·9. It is extremely difficult to fuse, but at a temperature below its fusing point it becomes pasty and may be welded—yet this welding is prevented by a very small portion of sulphur. Iron is eminently magnetic, and this is quite a characteristic property, no other metal possessing it, at least to the same extent. In combination with carbon it forms either cast iron or steel, which are also attracted by the magnet, as is well known. Exposed to a damp atmosphere iron becomes covered with rust, which is a hydrated oxyde of iron. Heated in the air it becomes covered with a crust, which is a combination of protoxyde and of peroxyde of iron; and if it be projected in small masses, such as iron filings for example, in a flame, it burns in a beautiful manner, and

with very vivid scintillations: no other metal presents this character. Hydrochloric acid dissolves iron rapidly, hydrogen gas being disengaged, and the result being protochloride of iron; moderately diluted *sulphuric acid* behaves in the same manner, with production of FeO,SO_3. Cold and dilute *nitric acid* dissolves iron and then produces protonitrate of iron, with evolution of nitrous acid fumes, but if heat be applied the solution contains pernitrate of iron. Moderately concentrated nitric acid attacks iron with great energy, ruddy fumes of hyponitric acid being abundantly produced and pernitrate of iron formed.

It however happens sometimes, that in treating iron by NO_5 the metal is dissolved without any gas being evolved, the hydrogen of the water combining with the nitrogen of the nitric acid to form nitrate of ammonia, whilst the nitric oxyde resulting from the partial deoxydisation of the nitric acid combines with the protonitrate of iron formed.

Sulphuric acid, if concentrated, and heated with iron, produces sulphate of iron and SO_2 is disengaged thus:

$$2\,SO_3,HO, + Fe = \\ FeO,SO_3 + 2HO + SO_2.$$

If, on the contrary, the acid be diluted, the water is decomposed, sulphate of iron is produced and hydrogen disengaged thus:

$$SO_3,HO + Fe = \\ FeO,SO_3 + H.$$

PROTOXYDE AND PROTOSALTS OF IRON.

(FeO.)

Protoxyde of iron in the pure and dry state is a black powder very rarely met with. Hydrated protoxyde of iron (FeO,HO) recently precipitated is *white*, but in the moist state it very soon absorbs oxygen, becomes *greenish*, and then *reddish-brown*; it is readily dissolved by acids. Its salts are colour-

less or of a bluish-green colour, and redden litmus paper: they are decomposed by a *red heat*; nitric acid converts them into persalts.

TESTS AND REACTIONS.

HS	*no precipitate*	in acid solutions, but the *neutral* solutions, *provided the acid be a weak one*, are incompletely precipitated, the precipitate being
	black.	
NH_4S		In neutral or in alkaline solutions a
	black	precipitate insoluble in excess, soluble in acids—in weak solutions *no precipitate*, but
	a dingy green	colour is produced.
KO	*White flocculent*	precipitate (FeO, HO) soon becoming
	dingy green	and then
	reddish-brown	on the filter, because it absorbs the oxygen of the air, and thus becomes converted into peroxyde. Salts of ammonia interfere with, or prevent the production of this precipitate.
NH_3	*White*	precipitate (FeO, HO) becoming
	dingy-green	and then
	reddish-brown	as with KO. Salts of ammonia completely prevent the formation of this precipitate
$K_3,2Cfy$	*Beautiful blue*	precipitate like Prussian blue (Fe_3, 2Cfy), insoluble in HCl, decomposed by alkalies.
NO_5		Moderately strong nitric acid poured upon a protosalt of iron imparts a
	brownish-black	colour to the liquid round the salt, if it is in the solid state, or if it is a *solution* of a protosalt of iron, the liquor becomes brownish black. An excess of NO_5 destroys this colour. (See Table V., Observation *h*, and the Dictionary of Reagents, Perchloride of iron.)
$AuCl_3$		Gold is reduced in the shape of a fine
	brownish-black	powder, in very dilute solutions.

The most characteristic tests for protosalts of iron are $K_3,2Cfy$—K_2Cfy, and NH_4S.

PEROXYDE AND PERSALTS OF IRON.

(Fe_2O_3)

Sesquioxyde or peroxyde of iron is generally of a fine reddish-brown colour, when in powder, and after ignition, it is almost black, and is then with difficulty soluble in acids. (See Table XIII, Observation *j*.) Its hydrate (Fe_2O_3,HO) is of a fine reddish-brown, orange, or drab-colour.

The persalts of iron are white, yellow, brown, or reddish-brown; their solution is always reddish-brown.

TESTS AND REACTIONS.

HS .	*White* . .	precipitate of *sulphur*, and the peroxyde is reduced to the state of protoxyde thus —$Fe_2O_3 + HS = 2FeO + HO + S$. (See Table IV, Observation *a'*, Table IX., Observation *h*; Table XV., Observations *c* and *d*)
NH_4S . . .	*Black* . .	precipitate (FeS) becoming
	brown . .	by exposure to the air, soluble in acids If only a minute trace of iron is present the liquor becomes
	greenish black.	
KO, or NH	. *reddish brown*	precipitate (Fe_2O_3,HO) bulky, insoluble in excess of the reagents, the presence of ammoniacal salts does not interfere.
K_2Cfy . . .	*Fine blue* .	colour and precipitate (Prussian blue) insoluble in acids, decomposed by KO into a
	reddish-brown	powder (Fe_2O_3)

When acetate of peroxyde of iron has been exposed to the heat of a steam bath for several hours, it undergoes a change in consequence of which K_2Cfy no longer produces a blue

precipitate in its solution, whilst concentrated HCl or NO_5 produce therein a brick-red granular precipitate entirely soluble in water.

$(2NaO),HO,PhO_5$.	*Whitish* .	precipitate, soluble in A,HO becoming
	Brown	in presence of a free alkali, and soluble in NH_3, provided there be an excess of $(2NaO), HO, PhO_5$. The solution has a
	Reddish brown	colour.
K_3,2Cfy . .	*No precipitate,* the liquor only becomes *greenish,* especially after dilution	
K,2CyS .	*Intense blood-red* colour.	
Infusion of galls .	*Bluish-black* colour (ink).	

The most characteristic tests for persalts of iron are NH_3, K_2Cfy—K_3(2Cfy)—K,CyS and Tincture of Galls.

LEAD.

(Pb.)

Lead is a very soft metal of a bluish colour, when newly cut or scraped, but the bright surface soon becomes tarnished by exposure to damp air, a grey coating of suboxyde being formed ; it soils the fingers, and leaves a streak like a pencil-mark when rubbed on paper. Its specific gravity is 11·445 when quite pure, but it is often met with in commerce of specific gravity 11·352 only, unlike the other metals its density appears to be diminished by hammering. It is malleable and ductile, but its tenacity is very feeble. Lead melts at a temperature rather above 600°, and at a white heat it boils and evaporates, but even at a less elevated temperature it volatilises in dense visible fumes. When melted and allowed to cool slowly, it crystallises in the form of quadrangular pyramids or in regular octahedrons. When immersed in *pure distilled water* and sheltered from the contact of the air, white spangles of hydrated oxyde of lead are formed.

Hydrochloric acid and dilute *sulphuric acid* have not much action upon lead, but hot concentrated SO_3,HO dissolves it, sulphate of lead being formed and SO_2 disengaged. *Nitric* acid dissolves it rapidly, and is in fact the best solvent of this metal. It is also soluble in $\bar{A},HO$.

PROTOXYDE AND PROTOSALTS OF LEAD.

(PbO)

Anhydrous protoxyde of lead (Litharge) has a yellow or drab-colour, the hydrate is white—calcination renders it red, and converts it into red lead (PbO_2,PbO). It melts at a red heat.

The salts of protoxyde of lead are colourless when their constituent acid is colourless. They have a sweet astringent taste and an acid reaction.

If the liquid or solution contains a great excess of a mineral acid, it must be diluted with water, or else the great excess of acid must be neutralised by either NH_3 or NaO,CO_2, otherwise HS will produce no precipitate.

When however the salts of lead are dissolved in a great quantity of HCl, a current of HS produces in the solution a *red precipitate* consisting of chloride and of sulphuret of lead.

TESTS AND REACTIONS.

HS . .	*Black* . . .	precipitate (PbS).
NH_4S .	*Black* . . .	precipitate (PbS).
KO or NH_3 .	*White* . . .	bulky precipitate (basic salt of lead), soluble in an excess of KO, especially with the help of heat.
	No precipitate	in solutions of acetate of lead, because basic acetate of lead is soluble in water.
KO,CO_2, or $NaO\,CO_2$	*White* . .	precipitate of (PbO,CO_2) soluble in KO.
HCl . . .	*White* . . .	heavy precipitate of chloride of lead (PbCl), soluble in a pretty large quantity of cold water, and in a

		smaller quantity of boiling water, from which it separates on cooling in shining crystals. The solubility of chloride of lead in water is much diminished by the presence of *chloride of calcium*, wherefore a solution of the latter salt produces a white precipitate in aqueous solutions of chloride of lead. (See Table XIV., Observation *a*; Table XXVIII. B, Observation *b*.)
KI . . .	*Yellow* .	precipitate of PbI in neutral solutions, soluble in a large excess of the reagent, especially by boiling but as the liquor cools, it is reprecipitated in beautiful golden spangles. (See Table II—B, Observation *h*).
SO_3, (dilute) or soluble sulphates	*White* .	precipitate of PbO,SO_3 insoluble, or nearly so, in water, slightly soluble in excess of dilute sulphuric acid, and of dilute nitric acid, soluble in the alkalies, in tartrate of ammonia, all of which solutions are rendered black by HS. Sulphate of lead is decomposed and converted into chloride of lead by HCl. (See Table I.—E, Observation *h*, Table XIII. Observation *d*.) If the solution be very acid and very dilute, a long time is often required for the precipitation of the sulphate, and the presence of ammoniacal salts (especially of NH_4O,SO_3) interfere with, or may even altogether prevent this precipitation.
KO,CrO_3 . .	*Yellow* . .	precipitate (PbO,CrO_3) insoluble in NO_5, soluble in KO. (See Table IV., Observation *q*.)
Zn.		A bar of metallic zinc immersed in solutions of salts of lead reduces the lead in beautiful spangles (lead tree).
Blowpipe+NaO,CO_2 . .		All the precipitates or salts of lead

MAGNESIUM.

(Mg.)

Magnesium is a metal of a white colour, looking almost like silver; it is malleable, melts at a red heat, and is not sensibly attacked by either hot or cold water, but exposed to a damp atmosphere it slowly oxydises, and is gradually converted into magnesia. The specific gravity of the metal, is 1·87. It is dissolved even in the cold, by the dilute acids, hydrogen being at the same time disengaged, but *nitric acid* whilst dissolving it, evolves nitrogen, and *concentrated* SO_3,HO, disengages sulphurous acid.

MAGNESIA (MgO) AND SALTS OF MAGNESIUM.

Pure oxyde of magnesium or magnesia is a white powder, almost insoluble in water, 5142 parts of which dissolve only one of magnesia, *hot water* dissolves a *still less* proportion. It combines with water to form hydrate of Magnesia (MgO, HO), but the combination is not attended with any evolution of heat, nor does it absorb carbonic acid so greedily as baryta, strontia, or lime. Most of the salts of magnesia are insoluble in water, but all are soluble in HCl, and are colourless. Their taste is bitter. All the salts of magnesia, except the sulphate and the phosphate, are decomposed by a red heat.

TESTS AND REACTIONS.

NH_4 . . .	*White* .	bulky precipitate in *neutral* solutions.
KO	*White* . .	precipitate (MgO, HO) *provided no ammoniacal salt* is present, the precipitate is immediately dissolved by all ammoniacal salts, especially by sal ammoniac. If, however, there be an excess of potash, and the solution be boiled, a white precipitate will

		appear. (See Table V., Observation *a*; Table VI., Observation *a*.) The precipitation of magnesia by KO is sometimes prevented by the presence of organic matter.
NaO,CO_2, or KO,CO_2	*White* . .	precipitate ($MgO, CO_2+(3MgO), HO$) in *neutral* solutions, which precipitate increases by boiling, and is immediately soluble in a solution of ammoniacal salts, and especially of sal ammoniac. (See Table VI., Observation *o*.) If the solution containing the magnesia is acid, a precipitate is produced by KO,CO_2 only by boiling the liquor.
$(2NaO),HO,PhO_5 + NH_3$	*White*.	precipitate ($(2MgO) NH_4O, PhO_5$,) which increases by boiling, is soluble in free acids, and is not interfered with by sal ammoniac. (See Table VI., Observation *m*, *n*. Table XX., Observation *b*)

Blowpipe. — Heated upon charcoal before the blowpipe, after being previously moistened with CoO,NO_5, and then heated again, a *faint rosy mass* is obtained (See Table XX., Observation *c*.)

The most characteristic tests for magnesia are $(2NaO),HO, PhO_5, +NH_3$—and the blowpipe with CoO,NO_5.

MALATES.—MALIC ACID.

($C_2H_4O_5$ 2HO or $\bar{M}$, 2HO)

Malic acid is a bibasic acid, that is to say, an acid which contains two equivalents of water which may be replaced by two equivalents of base to form a neutral salt. It is soluble in water, and the solution, evaporated to syrupy consistence, yields deliquescent crystals. It is also soluble in alcohol. The solution of malic acid has an

agreeable acid flavour, but it becomes mouldy by keeping. Ignited in the air it evolves an odour of burnt bread, or burnt sugar, like tartaric acid, and, submitted to distillation, pungent acid fumes of maleic acid are evolved, which condense in fine crystals on the cold parts of the tube, and a residue of *fumaric acid* is left.

Malic acid, like tartaric acid, prevents the precipitation of alumina, peroxyde of iron, and other oxydes.

Malic acid is transformed into oxalic acid by *nitric acid.*

Most malates are soluble in water, but all of them are soluble in nitric acid. Those of soda, potash, ammonia, and protoxyde of manganese are deliquescent.

TESTS AND REACTIONS.

CaCl + alcohol	*No precipitate.*	when poured *alone* in solutions of malic acid and of malates, either in the cold, or by boiling, but if *alcohol* be added, a
	white . .	precipitate of malate of lime ($CaO, HO,\bar{\bar{M}}$) is produced (See Table XXII. —A, Observation *k*).
PbO, A. . .	*White* .	precipitate $(2PbO)\bar{\bar{M}}$ which crystallises, when left at rest, into needles, which fuse in boiling water into a *viscid, transparent* mass.
AgO,NO_5. .	*White* . .	granular precipitate

MANGANATES.—MANGANIC ACID.

(MnO_3)

Manganic acid has never been obtained in the free state, and it is known only in combination with bases; taking manganate of potash (KO,MnO_3) as a type of this class of compounds, it is in the solid state a very dark green mass, which undergoes decomposition with the greatest facility, by parting with its oxygen when in contact with a great number of substances, and being thus resolved into Mn_2O_3, and KO. Its solution in water is intensely green,

and as it is decomposed by all organic substances, it cannot be filtered through paper. All acids, even the weakest, convert the green colour of the solution into a red or purple colour, because a portion of the manganate passes into the state of permanganate, which is red, a salt of protoxyde being at the same time formed; on applying heat, the solution is soon decolorised, because the permanganic acid is then decomposed. The acids in *ous*, such as sulphurous, phosphorous, hyposulphurous acids, &c., decolorise it also.

TESTS AND REACTIONS.

HCl .	*Reddens*	the green solution, but *decolorisation* soon supervenes, Cl being disengaged.
HS, or SO_2, or NH_4S	*Decolorisation.*	

The presence of manganese is easily detected by the usual reactions of the salts of manganese.

MANGANESE.

(Mn.)

Manganese is a metal of a greyish white colour, resembling white cast iron; it is hard, brittle, non-magnetic, and its specific gravity is 7·05. It has a great affinity for oxygen, and accordingly it becomes oxydised by exposure. When handled with a moist hand, it evolves an unpleasant garlicky odour, similar to that of the hydrogen which is disengaged when cast-iron is treated by an acid, and it slowly decomposes water. Metallic manganese must, therefore, be kept in naphtha, or in a glass tube, sealed at both ends. It is extremely difficult to fuse. It dissolves easily in the *dilute acids*, and hydrogen is evolved; the solution then contains a salt of protoxyde of manganese. *Nitric acid* dissolves manganese with evolution of nitric oxyde; the solution contains nitrate of protoxyde of manganese. Manganese forms with

oxygen a numerous series of compounds, namely, the protoxyde MnO, red oxyde Mn_3O_4, sesquioxyde Mn_2O_3, binoxyde, or peroxyde, MnO_2, manganic acid, MnO_3, permanganic acid, Mn_2O_7.

PROTOXYDE AND SALTS OF MANGANESE.

(MnO)

Protoxyde of manganese, though seldom met with in the pure state, is the base of nearly all the salts of manganese. The anhydrous protoxyde is of a greyish-green colour. The hydrate (MnO,HO) is white; but, both in the anhydrous state, and in that of hydrate, it gradually absorbs oxygen from the air, and becomes brown. After strong ignition, however, it may remain for a long time in the state of protoxyde. But the red oxyde (Mn_3O_4) being the most fixed oxyde of manganese, and the only one of these oxydes which is not altered by heat, is the state in which manganese is always quantitatively determined in chemical analysis.

The salts of manganese are either white, or of a pale pink colour. The latter colour is due to the presence of a trace of either cobalt, or of permanganic acid.

The neutral salts of manganese do not redden litmus paper. All the soluble salts of manganese, except the sulphate, are decomposed by a red heat.

TESTS AND REACTIONS.

HS . .	*Nothing.* . in neutral, nor in acid solutions of MnO.
NH_4S . .	*Flesh-coloured* precipitate, insoluble in an excess of the re-agent and in the alkalies; soluble in acids. The smallest trace of impurity modifies the colour of this precipitate which, by exposure to the air, gradually absorbs oxygen, so that the parts exposed to its contact

		soon become *brownish black* (See Table V., Observation *f*, and Table XVIII., Observation *d*.)
KO or NH_3	*White precipitate.*	(MnO,HO) at first, which by exposure becomes
	yellowish .	then
	brown . . .	and finally of a deep
	brownish-black . .	colour. (See Table V. Observation *p*.) The precipitate is not produced by NH_3, if ammoniacal salts are present.
$PbO_2 + NO_5$		If the solution in which Mn is suspected, be boiled with PbO_2, and dilute NO_5 (free from HCl), the liquor immediately assumes a fine reddish-purple colour, due to a formation of permanganic acid (Mn_2O_7); the smallest trace of Mn can be thus detected. This process is due to Mr Walter Crum.
$KO + KO,NO_5$.	. . . . .	Heated in a glass test-tube with a mixture of KO, and KO,NO_5, a
	green mass,	which is manganate of potash, (KO,MnO_3) is produced, which in dissolving in water forms a
	green. . .	solution which becomes *reddish-purple* by the addition of a *dilute acid* and is immediately *decolorized* by an addition of SO_2, or of an *organic matter* such as sugar, paper, starch, &c, a *brown* precipitate taking place at the same time.

Fused with NaO,CO_2, on a strip of platinum, in the oxydising flame of the blowpipe, especially if a little KO,NO be added, a mass of a *turquoise blue* colour is produced.

Blowpipe.—Heated with borax on a platinum wire in the oxydising flame of the blowpipe, a bead of an *amethyst* colour is obtained.—(See Table I.—B, Observation *o*.)

The most characteristic tests for manganese are NH_4S—

$PbO_2 + NO_5$—$NaO,CO_2 + KO,NO_5$—and the blowpipe with borax.

MECONIC ACID AND MECONATES.

($C_{14}\ HO_{11}$, 3HO or $\overline{Me}$, 3HO.)

The only source from which meconic acid is obtained is opium. Pure meconic acid is a tribasic acid, that is to say, it is an acid which takes three equivalents of base to produce neutral salts; it is in pearly spangles, is unctuous to the touch, and of a sweet astringent taste. In the crystallised state, meconic acid contains 9 equivalents of water, 6 of which it loses when dried at 212° Fahr., in which state it is dull and opaque. It is only sparingly soluble in cold water, but is very soluble in hot water, and in alcohol; the solution, however, undergoes decomposition by boiling, producing oxalic and comenic acids, and the solution becomes at first yellow, and then brown, carbonic acid being disengaged at the same time: a similar effect is produced when meconic acid is boiled with an excess of alkali. If boiled with HCl it is also converted into comenic acid with evolution of carbonic acid; but there is no production of colouring matter.

Meconic acid being a tribasic acid, as we just said, one or more equivalents of base is often replaced in meconates, by one or more equivalents of water.

All the soluble meconates which contain one or two equivalents of a fixed base, have an acid reaction; those which contain 3 equivalents of fixed base, have an alkaline reaction.

TESTS AND REACTIONS.

CaCl . . .	*White* .	precipitate ($CaO,2HO,\overline{Me}$) insoluble in sal-ammoniac.
Fe_2Cl_3 . . .	*Blood-red*	colour, but no precipitate; this blood red colour is not destroyed by solutions of HgCl (corrosive sublimate), and this is a distinctive character, since

the blood-red colour produced in the same circumstances by a sulpho-cyanide is immediately bleached by solution of corrosive sublimate.

It is true that the blood-red colour produced by Fe_2Cl_3 in solutions of the acetates of alkalies, is not destroyed either by solutions of corrosive sublimate; but, the reaction of the lime salts with meconic acid, is quite sufficient to distinguish it from acetic acid.

MERCURY.

(Me.)

Mercury is the only metal which is liquid at the ordinary temperature: it has a silver white colour, and is almost as brilliant as silver. At a temperature of —40° Fahr. it becomes solid, in which state it is soft and malleable. It boils at about 662° Fahr., and volatilises in the shape of very dense, transparent, and colourless vapours. Mercury can absorb a certain quantity of water and of air, from which it can be separated only by protracted boiling. Its specific gravity at 60° Fahr. is 13·56. Pure mercury becomes tarnished by exposure, but no oxyde appears to be formed on that account; if, however, it be heated in contact with the air, nearly to the boiling point, it gradually becomes oxydised, and converted into a beautiful dark red crystalline powder, which is peroxyde of mercury (HgO): at a dull red heat this red oxyde is decomposed, oxygen being evolved, and mercury reduced.

Nitric acid attacks mercury in the cold; and, if the mercury be in excess, even though heat be applied, subnitrate of mercury Hg_2O,NO_5, is produced, which crystallises as the liquor cools. But, if, on the contrary, the acid is in excess, then a solution of pernitrate of mercury (HgO,NO_5) is obtained. The addition of a large quantity of water to several of either of these neutral solutions, may produce a milkiness, which is due to the formation of insoluble basic compounds.

Dilute SO_3, has little, or, indeed, no action upon mercury; but, if the acid be concentrated, and heat applied, then, according as the acid or the metal is in excess, a basic sulphate (Hg_2O, SO_3), or a persulphate (HgO, SO_3) is produced, sulphurous acid being disengaged.

HCl, even in the concentrated state, has no action upon mercury out of the contact of the air; otherwise, Regnault has proved that water and perchloride of mercury are produced.

Mercury is completely soluble in aqua regia; the solution contains perchloride of mercury (HgCl, corrosive sublimate). A current of chlorine gas transforms mercury into perchloride, wherefore chlorine cannot be collected over mercury.

The solutions of the salts of mercury are colourless, and have a disagreeable metallic taste.

All the salts of mercury are decomposed or volatilised by a red heat.

SUBOXYDE AND SUBSALTS OF MERCURY.

(Hg_2, O.)

Suboxyde of mercury is decomposed and volatilised by heat in the form of a grey sublimate in which globules of mercury can be detected, either with the naked eye, or with the help of a magnifying-glass, especially by rubbing the sublimate against the sides of the glass-tube with a moistened glass-rod, or a piece of stiff paper. Suboxyde of mercury is a black, or dark olive powder, insoluble in water.

The salts of mercury, with excess of base, are often of a *yellowish* colour. Boiling with NO_5, or even with an excess of HCl converts them into persalts or into perchloride of mercury. (See Table XXVIII., Observation *f*.)

All the soluble salts of mercury have an acid reaction: they are all volatile, or decomposed by a moderate heat.

TESTS AND REACTIONS.

HS and NH_4S	*black* . .	precipitate (Hg_2S). If this sulphuret be carefully dried and heated in a glass test-tube closed at one end, a sublimate is produced in which globules of metallic mercury can be observed, and this reaction serves to distinguish Hg_2S from HgS, which, when heated in a like manner, produces also a sublimate of a similar nature, but in which no globules of mercury can be detected. Sulphuret of mercury is insoluble in dilute acids. (See Table XVII., Observation *a*,) and in an excess of NH_4S. (See Table XXIII., Observation *s*.)
KO . .	*black* . . .	precipitate (Hg_2O), insoluble in excess
NH_3 . .	*black* .	precipitate (Double salt of suboxide of mercury and ammonia), insoluble in excess.
HCl . . .	*White* . .	precipitate (Hg_2Cl), insoluble in cold HCl and in NO_5, but which is reduced by boiling with HCl into metallic mercury, HgCl being at the same time produced. If too much HCl is employed, and the boiling continued for some time, the whole of the *white* precipitate of the Hg_2Cl becomes converted into HgCl which is entirely dissolved. The precipitate produced by HCl immediately turns
NH_3 . . .	*black* . . .	when NH_3 is poured upon it. (See additions to Observation *b*, Table XIV.)
KI	*yellowish-green* . . .	precipitate (Hg_2I).
HCl and NH_3	*White* . .	precipitate (Hg_2Cl), which, treated by NH_3, turns *black*, owing to a separation of Hg_2O; dry Hg_2Cl becomes yellow by trituration.

Cu	*bright silver stain,*	If a clean and bright copper bar, or foil be rubbed with a rag moistened with the liquid under examination, a *bright silver stain,* which is reduced metallic mercury, is produced. The stain disappears by exposing the bar to a red heat.
KO,CrO_3 . .	*Brick red* . .	precipitate (Hg_2O,CrO_3).
NaO,CO_2	*grey sublimate*	If it be mixed in the dry state with NaO, CO_2, the mixture being slightly moistened, and then introduced into a small glass test-tube (see Table I. A, Observation *u*—*x*) and heated therein by means of the blow-pipe, a *grey sublimate* will be produced, reducible into globules of mercury, by rubbing with a glass-rod. (See Table XIII., Observation *a*)

The principal tests for the basic salts of mercury are HCl —NH_3—KO—and fusion with NaO,CO_2 before the blow-pipe.

PEROXIDE AND PERSALTS OF MERCURY.

Pure HgO is generally of a *red* or *orange-red colour,* and is in the state of a crystalline or shining powder. When obtained in the wet way, it is always amorphous, and yellow (HgO,3HO). When finely pulverised it is yellowish. Exposed for a long time to solar light its surface becomes black. Submitted to heat it turns *black,* but becomes red again on cooling. At a red heat it evolves pure oxygen, and condenses into metallic mercury. HgO is slightly soluble in water. It is a powerful oxydising agent, it detonates when heated in connection with sulphur, and it transforms chlorine gas into hypochlorous, and sulphurous into sulphuric acids.

The neutral salts of HgO are colourless.

TESTS AND REACTIONS.

HS . . .	*Black* . .	precipitate, which however passes through different shades of colour. At first the precipitate produced is
	white . . .	then it becomes
	yellow, orange, brownish-red,	as the liquid becomes more and more saturated, and finally it becomes
	black.	
NH_4S . . .	*black* . .	precipitate as with HS. The precipitate (HgS) is insoluble in an excess of reagent, and likewise in even boiling NO_5. (See Table XVII, Observation *a*) But it is readily soluble in aqua regia. (See Table IV., Observation *t*; also Table XXIII, Observation *s*.)
		If this sublimate be carefully dried and heated in a glass test-tube closed at one end, it sublimes, but without decomposition; whereas, Hg_2S—treated in the same manner—produces a sublimate in which globules of mercury can be readily detected, and hence a means of distinguishing HgS from Hg_2S.
KO . . .	*Yellow*, or *orange-yellow*	precipitate ($HgO,3HO$), insoluble in excess.
		If however the quantity of KO used is small, the precipitate is *brownish red*, and if ammoniacal salts are present the precipitate is white.
		If too small a quantity of KO is added, the precipitate is *reddish-brown*.
		Cyanide of mercury is not precipitated by KO.
NH_3 . . .	*White* . .	precipitate (composed of the salt of mercury and of amide of mercury).

		thus—$2HgCl+NH_3=HgNH_2, HgCl+HCl$ or NH_2Hg_2Cl. It is therefore a salt of ammonium, a portion of the hydrogen of which is replaced by mercury. Cyanide of mercury is not precipitated by NH_3.
KO,CrO_3 . .	*Fine red* .	precipitate (HgO,CrO_3).
KI	*Vermilion red*	precipitate (HgI), soluble in an excess of the reagent, and in an excess of the salt of mercury. Cyanide of mercury is not precipitated by KI.

Metallic copper reduces the salt of mercury, and a bright silvery stain is produced on the copper, which silvery stain disappears at a red heat.

Blowpipe.—Intimately mixed with NaO,CO_2, slightly moistened, and heated in a small glass tube, with the help of the blowpipe, a greyish sublimate is produced, in which metallic globules of mercury can be distinguished, by rubbing it with a glass rod. (See Table I.—A, Observation *u—x;* also Table XIII., Observation *h—i.*)

The most characteristic tests for the persalts of mercury are, NH_3—KO—KI—HS, and the blowpipe.

MOLYBDENUM.

Molybdenum is a white, rather brittle metal, very difficult to fuse. It is more generally met with in the state of a grey metallic powder, to which friction imparts lustre. It is not altered by exposure at the ordinary temperature; but, if it be then heated, it becomes converted first into brown oxyde (MoO), then into binoxyde (MoO_2), and, at a higher temperature, it forms white crystals of molybdic acid (MoO_3). Molybdenum is attacked violently by NO_5, and by aqua regia, and is converted by these acids into nitrate of molybdenum,

or into molybdic acid. Molybdenum is oxydised by concentrated SO_3,HO, with the help of heat, SO_2 being disengaged. HCl—dilute SO_3—PhO_5—HFl have no action on molybdenum.

TESTS FOR THE SALTS OF PROTOXYDE OF MOLYBDENUM.

KO or NH_3 . .	*Brown* . .	precipitate, insoluble in excess.
KO,CO_2 . .	*Brown* . .	precipitate, scarcely soluble in excess.
NH_4O,CO_2 .	*Brown* .	precipitate, very soluble in excess.
HS . . .	*Black* . .	precipitate, slowly formed.
NH_4S . . .	*Brown* . .	precipitate, soluble in excess.

TESTS FOR THE SALTS OF BINOXYDE OF MOLYBDENUM.

KO or NH_3 . . .	*Brown* .	precipitate, insoluble in excess.
KO,CO_2 or NaO,CO_2 or NH_4O,CO_2 .	*Brown* .	precipitate, soluble in excess.
HS . . .	*Brown* .	precipitate, slowly formed.
NH_4S . . .	*Brown* .	precipitate, soluble in excess; but the solution is precipitated yellow by HCl.

MOLYBDIC ACID.

(MoO_3)

Molybdic acid is white, sometimes in brilliant silky needles; ignition renders it yellow. It may be volatilised by heat; its sp. gr. is 3·5. It is sparingly soluble in water, 1 part of molybdic acid requiring 570 parts of water for its solution at the ordinary temperature. It is more soluble in boiling water. It combines with the alkalies, forming with them colourless crystallisable neutral and acid salts.

Molybdic acid is precipitated from its solutions by all acids, but an excess of the latter redissolves the precipitate at first produced.

The solutions of molybdates in acids, and diluted with water, become successively *blue*, *green*, and *black*, when a

bar of metallic zinc is plunged into them, and at last a rusty-looking deposit takes place. *Tin* behaves in the same manner.

Heated with *borax* upon a hook of platinum or upon charcoal in the oxydising flame of the blow-pipe, a colourless bead is produced; in the deoxydising flame, the bead is brown or reddish brown.

With *microcosmic salt* upon a hook of platinum or on charcoal, in the oxydising flame of the blowpipe, a colourless bead is obtained, in the deoxydising flame the bead is green.

NICKEL.

(Ni.)

Nickel is a white malleable metal with a slight tinge of grey. Its specific gravity is 8·8. This metal is almost as difficult to melt as manganese. It is magnetic, but it loses this property at about 660° Fahr. It is less oxydisable than iron, and is not altered at all by exposure, at the ordinary temperature.

HCl, moderately strong, dissolves nickel, especially with the help of heat, hydrogen gas being disengaged. Anhydrous chloride of nickel is in beautiful shining soft masses of a golden colour, resembling *aurum musivum*, and which in that state are very difficultly soluble in water; but if left exposed to moist air, they absorb water, lose their shining golden appearance, become green, and then dissolve readily in water.

It is also rapidly dissolved by NO_5.

Dilute SO_3 attacks it with more difficulty, but nevertheless dissolves it completely, especially with the help of heat.

Aqua regia dissolves it rapidly.

All the solutions of nickel in acids are green.

OXYDE AND SALTS OF NICKEL.

(NiO)

Dry oxyde of nickel is a grey powder. Its hydrate (NiO, HO) is apple-green. Both oxydes are soluble in acids. The

solutions of nickel have a green colour. The salts of nickel in the dry state are yellow. Anhydrous nickel is in beautiful and soft masses of a golden colour, resembling *aurum musivum*.

All the salts of nickel are decomposed by heat, though NiO,SO_3 resists its action for a long time. The neutral salts of nickel slightly redden litmus paper.

TESTS AND REACTIONS.

HS . . .	*Nothing* . .	in *acid* solutions. In neutral solutions, an insignificant
	Black . .	precipitate is produced, but in *alkaline* solutions, a
	Black . .	precipitate takes place.
NH_4S . . .	*Black* . .	precipitate, soluble in an excess of NH_4S, owing to which the superincumbent liquor has almost invariably a brownish black colour, which is due to a portion of NiS kept in solution. This brownish liquor, however, becomes clear by exposure, the NiS gradually depositing. NiS is somewhat soluble in HCl, even when the acid is dilute. If the NH_4S is *perfectly saturated* with HS, the nickel is completely precipitated
KO	*Apple-green*	precipitate (NiO,HO), insoluble in excess The presence of organic matter interferes with the production of this precipitate, but not so much as that of NH_3, or of ammoniacal salts. (See Table V., Observation *k*)
NH_3.		produces, at first, a precipitate of an
	Apple-green. .	colour (NiO,HO); but an excess of NH_3 redissolves it, and the liquor assumes a
	Beautiful blue	colour, which very much resembles that produced under the same circumstances in salts of copper. The presence of ammoniacal salts,

and of free acids, prevent the production of this apple-green precipitate; but the blue colour of the liquor always follows the addition of an excess of NH_3. (See Table XXXI, Observation *a*.)

KCy . . . *Greenish white* precipitate ($NiCy$); soluble in an excess, and the solution has then a brownish colour; but the white precipitate is reproduced by the addition of an acid.

The most characteristic tests for nickel are NH_3—NH_4S—KCy.

NITRATES—NITRIC ACID.

(NO_5)

Nitric acid is scarcely known in the anhydrous form. In a free state it forms with water, when pure, a colourless corrosive solution, possessing strongly acid and oxydising properties. In the concentrated state it emits fumes in the air, and destroys organic substances, most of which are thus resolved into carbonic acid and water; those which contain nitrogen are stained a deep yellow by the concentrated acid. All metallic oxydes, except those of tin, antimony, tellurium, and tungsten, are dissolved by it. All the neutral nitrates are soluble in water; a few oxydes, such as suboxide of copper (Cu_2O), and oxyde of bismuth, form basic nitrates insoluble therein. All nitrates are decomposed by exposure to a red heat: those the base of which is an alkali yield a mixture of oxygen and nitrogen; the others yield a mixture of oxygen, nitrous or hyponitric acids, a residue being left. Nitrate of ammonia is converted by a red heat into *Nitrous oxyde* and water. When nitrates are thrown upon red hot charcoal, or heated in a glass test tube with a little paper or charcoal, deflagration is produced; and if a pulverized nitrate be thus heated with a little cyanide of potassium upon a piece of

platinum foil, deflagration takes place attended with a detonation.

The tests for nitrates are in general less delicate than for the other acids, wherefore it is always advisable to concentrate the solution under examination before submitting it to the action of reagents. If the liquid has an acid reaction it is well to neutralize it with potash before evaporating it.

TESTS AND REACTIONS.

Solution of Indigo + SO_3,HO		If solution of indigo be added to the liquor, so as to impart a slight blue tinge to it, then a few drops of SO_3,HO, and if the whole be boiled, the blue colour, at first, becomes yellowish, and shortly after the solution is entirely
	Decolorized.	This effect is produced, also, by free *chlorine*, the *chlorates*, and various other substances.
Copper-filings + SO_3,HO		Heated with copper-filings, and one-fourth of its bulk of concentrated sulphuric acid, in an open test-tube,
	Ruddy fumes	of nitrous acid are evolved, especially by adding water (See Table I.—C Observation *v*.) Care must be taken that the SO_3,HO employed contains no nitric acid, since, in that case, the test would, of course, be fallacious.
FeO,SO_3 + SO_3,HO		If to the solution under examination about one-half of its bulk of concentrated SO_3,HO is added, and then after cooling, a crystal of FeO,SO
	a brownish or greenish black	colour will be perceived all round th crystal. This colour disappears b shaking, but more readily by heat ing the liquor. (See Dictionary Reagents, Art. Proto-sulphate Iron.)

Pen-shavings. Free nitric acid is detected by boiling th solution, in which its presence is suspected, with a few pen nail-scrapings, which will then become yellow, either throug

out, or if the quantity of nitric acid present is very small on the edges only. This reaction is quite characteristic of the presence of *free* nitric acid, since even the acid salts do not exhibit this reaction. (See Table VII., Observation *u'*,).

The most characteristic tests of the presence of nitric acid in the state of nitrate are, the reaction with *copper-filings* and *sulphuric acid*, the reaction with *sulphuric acid*, and FeO,SO_3.

Nitrates are also distinguished from other inorganic acids by their producing no precipitate with BaCl, nor with AgO, NO_5. Chlorates, however, yield no precipitate either with these reagents, but they do not behave with copper-filings, and with FeO,SO_3 as nitrates do.

Brucine and *Morphine* are also excellent and delicate tests for nitric acid and nitrates. (See in the Dictionary of Reagents and Tests, Art Brucine and Morphine.)

The yellow colour imparted by boiling with *pen* or *nail scrapings* is quite characteristic of the presence of *free* NO_5.

NITROGEN.

(N.)

Nitrogen is a permanent, colourless, tasteless, and inodorous gas, which cannot support respiration or combustion, but is not otherwise deleterious. It is a little lighter than common air, its specific gravity being 0·973, and therefore 100 cubic inches of it weigh 30·14 grains at the ordinary temperature and pressure. It is less soluble in water than oxygen, this menstruum dissolving only $\frac{16}{1000}$ part of its bulk. It is not inflammable, is not absorbed by solution of KO, does not render lime-water turbid, and in fact it is the very negative character of its behaviour towards reagents which serves to identify it.

PROTOXYDE OF NITROGEN.

(NO.)

Protoxyde of nitrogen at the ordinary temperature is a colourless, transparent, and almost inodorous gas, of a sweet or slightly sugary taste. Its specific gravity is 1·525, and consequently 100 cubic inches weigh 47·29 grains. It supports combustion with nearly as much energy as oxygen, but is much more soluble in water than the latter, as shown in Table XXXIII., Observ. *i.* It can be breathed for a short time without inconvenience, and from its intoxicating or exhilarating properties, it has been called *laughing gas.* Water dissolves half, and alcohol one and a half times its bulk of this gas.

OSMIUM.

(Os.)

Osmium, when in the compact state, is a metal of a bluish-white colour, somewhat malleable, yet easily reduced to powder. Its specific gravity is 10—it is neither volatile nor fusible, but when heated in contact with the air, it is converted into osmic acid, easily recognisable by its peculiar disagreeable odour, which resembles that of chloride of sulphur.

At the ordinary temperature, it is not oxydised. It is dissolved by concentrated NO_5, ruddy fumes of hyponitric acid being disengaged, the result being osmic acid (OsO_4) ; yet the solution is slow, and if heat be applied, the two acids evaporate together. Aqua regia dissolves it more easily and rapidly, osmic acid (OsO_4) being also produced. *Fuming nitric acid* is its best solvent; yet after exposure to a very high temperature in close vessels, osmium can no longer be dissolved by these acids, and it is then necessary to fuse it with nitrate

of potash, or with some of the alkalies by which it becomes converted into osmiate of alkali.* Osmium forms several combinations with oxygen, the only important ones being the deutoxyde or binoxyde (OsO_2), and osmic acid, (OsO_4.)

TESTS FOR SALTS OF DEUTOXYDE OF OSMIUM.

KO	*Black* . .	precipitate, requiring some time for its production which is promoted by boiling.
NH_3 or KO,CO_2 .	*Brown* . .	precipitate, after some time.
NH_4Cl . .	*Red* . . .	precipitate.
SnCl . . .	*Brown* . .	precipitate.
Hg_2O,NO_5 .	*Yellowish-white* . .	precipitate.
HS	*Yellowish-brown* .	precipitate after some time.
NH_4S . . .	*Yellowish-brown* .	precipitate insoluble in excess.

OSMIC ACID.

(OsO_4.)

Osmic acid is in colourless, brilliant, flexible, and regular prisms, which melt and boil at a heat below 212° Fahr. The vapour of osmic acid has a most disagreeable odour, resembling that of chloride of sulphur, and is extremely irritating and deleterious. It has a powerful action on the skin, to which it gives a dartrous appearance, and stains it black; it is very soluble, though very slowly in water: several metals, such as zinc, tin, iron, and copper, reduce it completely, osmium being deposited. Osmic acid is a very weak acid which does not redden litmus paper, nor even decompose carbonates, but it dissolves in the alkalies with which it forms salts, which, when the base is in excess, are always of a brown colour; all osmiates are decomposed by boiling, osmic acid being disengaged.

TESTS AND REACTIONS.

FeO,SO_3 . .	*Deep black* .	precipitate, and the osmic acid is reduced.
NaO,SO_2 .	*Deep violet blue* . .	tinge, and
	Black precipitate .	of metallic osmium (See Table XVII. Observation *h*.)
HS	*Brownish-black* .	precipitate
NH_4S . . .	*Brownish-black* . .	precipitate, soluble in excess.
NO_5 or HCl .	*Odour of decayed radish.*	

The disagreeable odour of free OsO_4 is quite characteristic.

OXALATES.—OXALIC ACID.

(C_2O_3,HO or $\overline{O}$,HO)

It is very probable that oxalic acid is a bibasic acid, and that its equivalent is double that which is given in the above symbol, namely C_4O_6, 2HO, which would account for the compounds called acid oxalates, and double oxalates.

Oxalic acid forms rhombic colourless crystals, or four-sided prisms, very soluble in water, and often emitting a peculiar crackling noise whilst dissolving. The crystals by exposure crumble into a white powder on account of their losing a portion of their water. They are soluble also in alcohol. When the crystals are rapidly heated, a portion of the acid sublimes without decomposition, the other portion being thereby converted into carbonic acid (CO_2), carbonic oxide (CO), and formic acid (C_2HO_3, HO or $\overline{Fo}O_3$,HO). Whilst volatilising, the fumes evolved are very pungent and irritating. The destructive distillation of $\overline{O}$,HO, is not attended with

any charring, and this, amongst others, is a distinctive character, since most other organic acids are charred by a red heat.

The oxalates of alkalies only are soluble in water, the other oxalates are insoluble therein, except *oxalate of tin*, of *chrome* of *alumina*, and of *iron*. The oxalates of earths are not soluble in $\overline{A}$,HO. But all oxalates are soluble in the strong acids. All oxalates are destroyed by heat; those of the earths and of the alkalies leave a residue which is a carbonate of the base, those of the metallic oxydes leave a residue which is a metal or an oxyde.

The oxalates of earths are precipitated from their acid solutions by KO,$\overline{A}$,HO, the reagent being decomposed by the free acid, and replaced by acetic acid, in which the oxalates of earths are insoluble.

TESTS AND REACTIONS.

BaCl . . .	*White* . .	precipitate (BaO,$\overline{O}$); sparingly soluble in water, but soluble in HCl and in NO_5; wherefore the solution tested should be neutral, or else no precipitate will be produced. (See Table VII., Observation *b'*, Table XXII. A, Observation *h*.)
CaCl or CaO,SO_3 or Lime-water .	. . . *White* . .	All the soluble salts of lime produce a precipitate, insoluble in water, and only sparingly soluble in very dilute HCl, and in NO_5; but readily soluble in these acids when more concentrated, wherefore the solution should be neutralized by NH_3, which greatly promotes besides the agglomeration of the precipitate. (See Table XXII.—A., Observation *o*.)
AgO,NO_5 . .	*White* . .	precipitate; soluble in NO_5, and in NH_3. (Table VII., Observation *c'*,) this precipitate fulminates slightly when heated, and leaves no residue.
CuO,SO_3 .	*Greenish white* .	precipitate; sparingly soluble in HCl.

SO_3,HO . . . . Heated in a glass tube, or small retort, with concentrated SO_3,HO carbonic acid and carbonic oxyde are copiously disengaged; the latter being inflamed, burns with a

Blue flame. (See Table I.,—C., Observations *g*, *i*)

The sulphuric acid takes all the water, and as anhydrous oxalic acid cannot exist in a separate or free state, it is decomposed. The reaction is very simple, and is as follows:—

$$\overbrace{C_2O_3,HO}^{\text{Oxalic acid}} + SO_3,HO = CO + CO_2 + SO_3,2HO.$$

It is better, instead of using SO_3,HO alone, to add some finely levigated MnO_2 to the compound or liquor under examination, and then a small quantity of SO_3,HO; a disengagement tube is then fixed to the tube, and the gas evolved by boiling is collected as described in the next paragraph. The rationale of the reaction is as follows:—

$$\overbrace{C_2O_3}^{\text{Oxalic acid}}HO + MnO_2 + SO_3 = MnO,SO_3 + 2CO_2.$$

The MnO_2 should be first washed with dilute NO_5, in order to remove the earthy carbonates which may be and often are present, and which would altogether interfere with the accuracy of the test.

When the quantity under examination is small, and the compound is a liquid, the carbonic acid, and the aqueous or acid steam evolved at the same time interfere with the inflaming of the gas. The best method perhaps of operating, whether the substance is in the liquid or in the solid state, is to transfer a portion of it (which, if liquid, should be very concentrated) to a glass test-tube, or small flask provided with a perforated cork, and having added the concentrated acid, and applied heat, to collect into the tube B the gas thus evolved, and which may be readily inflamed. When the

substance to be examined, as just said, is a liquid, not only it should be very concentrated, but the operator should also defer collecting the gas *until the disengagement becomes manifest.* The burning of the gas with a blue flame is more apparent if the operator stands against daylight, and if, before he inflames the gas, he shakes it with a small fragment of KO, rapidly introduced into the tube closed with the thumb. After shaking, the thumb is sucked in by the absorption of the CO_2, and no other gas is left in the tube than CO, which, on removing the thumb, may be inflamed, and which then burns with its characteristic blue flame. (See Table VII., Observation *x*.)

The principal and most characteristic tests for $\overline{O}$, HO, are the soluble salts of lime, and its reaction when heated with concentrated SO_3,HO.

The greenish-white precipitate produced by CuO,SO_3, *in acid solutions,* is also very characteristic, since none of the mineral acids, nor citric, tartaric, or acetic acids are precipitated by that reagent.

But perhaps the most characteristic test for oxalic acid or oxalates in solutions, is that which is produced with solution of CaO,SO_3, since none of the inorganic acids which sulphuric acid displaces present this reaction even in presence of other salts, except those of baryta and strontia

OXYGEN.

(O.)

Oxygen in the free or uncombined state is a permanent, colourless, tasteless, inodorous gas, of a specific gravity, 1.1057, and therefore 100 cubic inches of it weigh 34·29. At the ordinary temperature and pressure, it is sparingly soluble in water, 1 part of oxygen requiring 27 of water for its solution. It is the best supporter of combustion, and this is in fact one of its most characteristic properties, oxygen and nitrous oxyde

being the only gases which can rekindle a glowing splinter plunged into them, but nitrous oxyde does this much less energetically, and the combustion is less vivid. Mixed with hydrogen, or with gases of which hydrogen is one of the constituents, a violent detonation is produced when the mixture is inflamed.

A bubble of oxygen, admitted in a jar or vessel containing NO_2, at once produces ruddy fumes of nitrous acid. It is absorbed by phosphorus, by the alkaline sulphurets, by an ammoniacal solution of protosulphite or of protochloride of copper, by the solution of a mixture of 2 parts of caustic potash and 5 parts of crystallised protosulphate of iron, by a solution of extract of logwood, by a concentrated solution of potash and pyrogallic acid, all these substances are employed for the determination of the amount of oxygen contained in common air.

PEROXYDE.	
PERSALTS.	SEE THE METAL, OR THE METALLOID.
PROTOSALTS.	
PROTOXYDE.	

PALLADIUM.

(Pd.)

Palladium is a white metal, almost like silver, but with a yellowish-grey hue, the fusion of which requires an exceedingly high temperature. Its specific gravity is 11·8. At a white heat it may be welded; it is very malleable and ductile. There are two oxydes of Palladium, the protoxyde PdO, and the binoxyde PdO_2, but this metal does not combine with oxygen in a direct way, the latter oxyde is not known in an uncombined state. The protoxyde is produced by fusing the metal with nitrate of potash. The metal is not attacked by SO_3,HO,

but it is easily dissolved by NO_5, with the help of heat, and it unites directly with sulphur, with chlorine, and with silver.

PROTOXYDE AND SALTS OF PROTOXYDE OF PALLADIUM.

Protoxyde of palladium in the anhydrous state is a dark grey powder, which loses all its oxygen by exposure to a high temperature; the hydrated oxyde is dark brown, and is easily dissolved by the dilute acids. The salts of protoxyde of palladium have a brownish-red colour.

TESTS AND REACTIONS.

HgCy . . .	*White* . .	precipitate. (See Table XXIV., Observation *a*.)
KO . . .	*Brown* . .	precipitate; soluble in an excess of the reagent.
HS or NH_4S . .	*Black* . .	precipitate; insoluble in excess of NH_4S.
NH_3 . . .	*Decolorisation*,	at least in solutions of PdCl. (See Table XVII., Observation *g*.)

PERMANGANIC ACID.

(Mn_2O_7)

Permanganic acid is a brown crystalline mass, very soluble in water, but most easily decomposed, a temperature of 60 or 70 Fahr. being sufficient to resolve it into oxygen and deutoxyde of manganese. This decomposition is effected in the cold by the contact of sugar, gum, paper, dust, and in fact of all organic substances, and by hydrogen, hydrocarbons, the acids in *ous* and other deoxydising agents. It forms with the alkalies, salts, which are soluble in water, and the solution has a

splendid purple colour. In the solid state those salts are almost black. The intense purple or crimson colour of the solution of permanganates, the facility with which they are decomposed into oxygen and peroxyde of manganese, *the copious disengagement of chlorine gas evolved when treated by* HCl, and their behaviour before the blowpipe, which is the same as for all the salts of manganese, are perfectly characteristic tests.

PHOSPHORIC ACID AND PHOSPHATES.

(PhO_5, 3HO.)

Phosphoric acid combines with water in three different proportions; these combinations have each a different capacity of saturation, and are known respectively under the names of *metaphosphoric, pyrophosphoric* and *phosphoric acid*, or *monobasic, bibasic,* and *tribasic phosphoric acid,* because the water of combination of each of these acids may be replaced by a corresponding number of equivalents of bases to produce a neutral salt. The monobasic (PhO_5,HO), and the bibasic phosphoric acid (PhO_5,2HO), in contact with water, gradually become converted into tribasic phosphoric acid (PhO_5,3HO). Each of these forms of phosphoric behave differently with some reagents, as will be indicated presently.

TRIBASIC PHOSPHORIC acid is met with in beautiful perfectly transparent and colourless crystals, known as *glacial phosphoric acid.* These crystals, however, are very deliquescent, and produce, by exposure, a syrupy liquid which has an acid, but not a caustic taste. When exposed to a very high temperature, phosphoric acid evaporates completely, but at a lower heat it becomes converted into bibasic or monobasic phosphoric acid, according as either one or two equivalents of water are expelled. The metaphosphates and pyrophosphates are reconverted into tribasic or ordinary phosphates by boiling

them with acids, and especially with concentrated SO_3,HO or by fusion with NaO,CO_2.

It must not be fused in glass nor in porcelain vessels, because they would be corroded by it.

Anhydrous phosphoric acid is a white powder which has such an affinity for water that, when put in contact with it, a hissing noise is heard as when red hot iron is immersed in that liquid.

The neutral or basic phosphates of alkalies are soluble in water, and their solution has an alkaline reaction on litmus paper; the other neutral phosphates are almost or quite insoluble in water. (See Table XIX., Observation *d.*)

TESTS AND REACTIONS.

$BaCl$		In solutions of neutral or basic phosphates,
	White	precipitate $(2BaO)PhO_5$, or $(3BaO)PhO_5$, according as the phosphoric acid is bibasic or tribasic; insoluble in HCl, and in NO_5, slightly soluble in NH_4Cl.
$CaCl$, or CaO,SO_3 . . .		In solutions of neutral or basic phosphates,
	White .	precipitate ($2CaO$, or $3CaO,(PhO_5)$; soluble in HCl, in NO_5, and even in $\bar{A},HO$.
MgO,SO_3 + NH_3		And other soluble salts of magnesia, produce a
	White crystal.	precipitate, which takes some time to appear, but which is produced *immediately* by further adding NH_3 to the solution. This white precipitate is a double basic phosphate of magnesia and ammonia $(2MgO)NH_4O,PhO_5)$; insoluble in NH_3, and in salts of NH_3, but soluble in free acids, even in $\bar{A},HO$. Stirring promotes the complete formation of this

		precipitate. (Table XI., Observation *m*, Table XIX., Observation *d*; Table XXII.,—A., Observation *n*.)
$PbO,\overline{A}$ . . .	*White* . .	precipitate ($3PbO, PhO_5$); soluble in NO_5, sparingly so in $\overline{A},HO$ · this precipitate, dried and fused on charcoal before the oxydising flame of the blowpipe, yields a bead which crystallises on cooling.
AgO,NO_5		In solutions of neutral or basic phosphates, a
	Lemon yellow	precipitate ($(3AgO),PhO_5$), very readily soluble in NO_5, and in NH_3. This precipitate is quite characteristic, since *arsenious* and *silicic acids* are the only acids which yield a yellow precipitate with this reagent. (See Table VII., Observation *c'*.)

The neutral phosphate of soda ($(2NaO), HO, PhO_5$), or ordinary phosphate of soda, being a tribasic phosphate in which 1 equivalent of oxyde is replaced by 1 equivalent of basic water, yields of course this yellow precipitate of tribasic phosphate of silver ($(3AgO) PhO_5$); its solution before testing with AgO,NO_5 has a feebly *alkaline* reaction, but after testing it therewith, the supernatant liquor has an *acid* reaction: this effect may be explained by the following equation—

$$\overbrace{(2NaO),HO,PhO_5}^{\text{Tribasic phosphate of soda.}} + 3(AgO,NO_5) = \underbrace{3AgO,PhO_5}_{\text{Tribasic phosphate of silver (yellow).}} + 2(NaO,NO_5) + NO_5,HO$$

By which it is seen that 1 equivalent of NO_5 has been liberated, and hence the liquor which was alkaline before, becomes acid after testing with AgO,NO_5.

When, however, common neutral phosphate of soda ($(2NaO), HO,PhO_5$) is exposed to a red heat, it loses its equivalent of water, and becomes thereby converted into pyrophosphate of

soda, or bibasic phosphate of soda $(2NaO), PhO_5$), and the solution when tested with AgO,NO_5 yields now a precipitate which, instead of being YELLOW as before, is WHITE. This apparently anomalous reaction is not extraordinary, since pyrophosphoric acid $(PhO_5, 2HO)$, requiring only 2 equivalents of base to form neutral salts, is an acid essentially different from tribasic phosphoric acid $(PhO_5, 3HO)$, which requires 3 equivalents of base to form such salts, and therefore the different precipitate which it forms with AgO,NO_5 is thus easily accounted for. The liquid resting upon the precipitate produced by AgO,NO_5 in solutions of pyrophosphate of soda is *neutral*. This reaction may be explained by the following equation—

$$\overbrace{(2NaO),PhO_5}^{\text{Bibasic phosphate of soda}} + 2(AgO,NO_5) =$$
$$\underbrace{(2AgO),PhO_5}_{\text{Bibasic phosphate of silver (white)}} + 2(NaO,NO_5).$$

If the liquor to be tested is alkaline, the best way, perhaps, of operating consists in slightly supersaturating it with a little NO_5, adding AgO,NO_5, and then pouring in gradually and carefully some NH_3 until the solution is nearly neutral, when a *yellow* or a *white* precipitate of tribasic, or of bibasic phosphate of silver, will be produced.

If the solution contains HCl, add a few drops of NO_5, and then an excess of AgO,NO_5, until it no longer produces a white precipitate (AgCl), boil, filter, and test the filtrate again with AgO,NO_5, to make sure that all the chlorine is removed, and that the liquor contains an excess of AgO,NO_5; add then to the clear filtrate some dilute NH_3, until the liquor is very nearly neutralised, when the *characteristic yellow precipitate* of tribasic (or *white precipitate* of bibasic) phosphate of silver will be produced. This precipitate being immediately soluble in a slight excess of NH_3, and of NO_5, care must be taken that the liquor contains no excess of either; or if such an

excess exists, neutralise it by adding NH_3 or NO_5 thereto as the case may be. (See Table VII., Observations *y*, *z*.)

Fe_2Cl_3 + $KO,\bar{A}$	. . . .	produces, in all solutions of phosphates to which a little HCl is added only to *acid reaction*, and then some $KO,\bar{A}$, a
	Yellowish-white gelatinous .	precipitate of perphosphate of iron $(2Fe_2O_3),(3PhO_5),3HO)$. Take care not to use more than a drop of Fe_2Cl_3, because an excess of this reagent would produce acetate of iron, which would impart a reddish-brown colour to the liquor. An excess of Fe_2Cl_3 may afterwards be added, which will cause the liquor above the precipitate to appear red, an excess of $KO,\bar{A}$ being at the same time added, until the liquor becomes of a very dark brownish-red colour, and the precipitate at first produced disappears. The liquor so prepared, after being boiled for a few moments, will yield a dark reddish-brown precipitate, and the supernatant liquor will be clear. The dark reddish-brown precipitate produced contains all the phosphoric acid, and is a mixture of perphosphate of iron, and of hydrated sesquioxyde of iron. (See Table IX., Observation *n*; Table XI, Observation *m*, Table XIX., Observations *a*, *d*.)
NH_4O,MoO_3 .	*Bright greenish yellow*.	colour, and precipitate in the somewhat concentrated hot and boiling nitric or hydrochloric acid solutions of phosphates. The yellow colour becomes fainter as the liquor cools.

Another excellent and characteristic test for phosphoric acid consists in heating in a small glass test-tube a portion of the well-pulverized compound, until all moisture is

expelled, then dropping upon the mass in the tube a small globule of metallic potassium, well dried with a towel, or piece of bibulous paper, and heating cautiously the whole to a red heat, for which the flame of an ordinary spirit-lamp is sufficient. When cold, empty the tube upon a slab, or porcelain crucible cover, and breathe over it; if phosphoric acid, or a phosphate was present, a strong odour of onion, or garlic, due to phosphuretted hydrogen, will be immediately perceived. *Do not add too much potassium,* as otherwise it will not be possible to empty the tube, in which case, however, the lower part may be cut off, and on breathing upon its contents the odour will be perceived. The reader will do well also to consult Table XVIII. additions to Observation *i*, in which various other processes are given for the detection of phosphoric acid.

Blowpipe.—Moistened with concentrated SO_3,HO, and heated in the oxydising or exterior flame of the blowpipe, phosphates impart a greenish blue colour to the flame.

The most conclusive tests for Phosphoric acid and phosphates, are MgO,SO_3+NH_3. The crystallised bead obtained by fusing with $PbO,\bar{A}$ before the blowpipe. The precipitate obtained by $PbO,\bar{A}$. The yellow precipitate produced by AgO,NO_5. The reaction with $Fe_2Cl_3+KO,\bar{A}$, and last but not least, the reaction with NH_4O,MoO_3.

PHOSPHOROUS ACID AND PHOSPHITES.

(PhO_3).

Anhydrous phosphorous acid is white, solid and volatile; it is very soluble in water; it absorbs slowly the oxygen from the air, and becomes converted into phosphoric acid.

The phosphites of alkalies only are soluble in water, but all phosphites are soluble in acids.

TESTS AND REACTIONS.

HgCl . . .	*White* . .	precipitate (Hg_2Cl); in neutral solutions; the precipitate is insoluble in HCl.
SO_2 . . .	*Milkiness* .	which is due to a deposit of sulphur.
$AgO,NO_5 + NH_3$ . .	*Blackish-brown* .	powder of metallic silver. It is best to add *an excess* of NH_3 first, the liquor assumes first a *brownish*, then a *brownish-red colour*, becoming *purple, grey,* and then a *black* deposit is produced; the reaction is hastened by boiling.
CaCl . . .	*White* .	precipitate; soluble in $\overline{A}$, HO. (See Table XXI., Observation *h*; Table XXII.,—A, Observation *j*.)

Blowpipe.—Heated before the blowpipe, the first impression of the heat causes phosphites to burn like phosphorus.

PHOSPHORUS.

(Ph.)

At the ordinary temperature common phosphorus is a transparent and tasteless substance, of a faint straw-colour, waxy consistence, and alliaceous odour—but $\frac{1}{500}$ of sulphur is sufficient to render it brittle; its specific gravity is 1·77; it melts at 108° Fahr., and boils at 550°. It is insoluble in water, and is usually kept immersed in that liquid, but it dissolves in oils, in petroleum; its best solvent, however, is bisulphuret of carbon. Phosphorus is luminous in the dark. It is exceedingly inflammable, and it burns with a most brilliant flame, and an emission of very thick white fumes, which consist of anhydrous phosphoric acid. Friction or a blow is sufficient to inflame it, and it frequently even bursts spontaneously into flame. In close vessels it may be volatilized and distilled by heat. Although the affinity of phosphorus for oxygen is very great, this gas does not act upon it when perfectly pure and

dry; but otherwise, phosphorus combines slowly with the oxygen of the atmosphere at the ordinary temperature; this slow oxydisation is prevented, however, by the presence of a small quantity of the vapour of ether, of naphtha, of essential oils, or of olefiant gas.

There is another kind of phosphorus in an allotropic or amorphous state, which is in black, brown, buff, or red masses, and which is inflammable only at a much higher temperature.

All these properties are so peculiar, that phosphorus cannot be mistaken for anything else.

PLATINUM.

(Pt)

Platinum is a metal, the colour of which is between that of silver and of steel. It is very malleable and ductile; it may acquire a fine polish, and is inodorous, tasteless, and inalterable in the air. It is softer than silver and iron, but its hardness is increased by traces of iridium: it is harder than copper. Of all metals, platinum is that which suffers least dilatation by heat. Its specific gravity is from 21·47 to 21·53. It is infusible at the highest heat of a blast furnace. It is only the heat of a strong electric battery, or of the oxyhydrogen blowpipe that can melt it. Yet at a white heat it becomes softened, and may be welded and forged. It is not oxydised by exposure at any temperature, and the ordinary acids have no action upon it.

Nitric acid has no action on pure platinum; but it is dissolved by that acid when alloyed with a sufficient quantity of silver and gold.

Its true solvent is *aqua regia.*

Sulphuric and *hydrochloric acid* have no action.

Potash and *lithia* attack platinum, and so does *soda,* but in a less degree.

Bisulphate of potash has also an action upon platinum; and

a mixture of *nitre and of potash* attacks platinum powerfully. (*See* what has been said at page vii. of the "Chemical Atlas" on the treatment of platinum utensils.)

When platinum is precipitated from its solutions, it is sometimes in the state of a black powder, called *platinum black* or *platinum mohr*. This powder has the property of condensing gases, and of slowly burning combustible substances.

OXYDE AND SALTS OF PLATINUM.

(PtO_2.)

Hydrated binoxide of platinum is a reddish-brown powder. The salts of platinum are of a dark orange-yellow, or reddish-brown colour. The aqueous solution of $PtCl_2$ resembles much in colour that of neutral Fe_2Cl_3. The neutral and soluble salts of platinum redden litmus paper. All the salts of platinum are decomposed by a red heat.

TESTS AND REACTIONS.

HS	*Black* .	precipitate (PtS_2), in acids and neutral solutions, but nothing in alkaline solutions, because PtS_2 is soluble in alkaline sulphurets. It is insoluble in the simple acids, but it dissolves in aqua regia. In acid, or in neutral solutions, of bichloride of platinum, HS at first renders the liquid brown, and it is only after some time that a black precipitate settles, the liquor remaining brown
NH_4S . . .	*Black* .	precipitate (PtS_2); soluble in a pretty large excess of the reagent, especially if it contains an excess of S the precipitate is soluble in aqua regia.
KO or NH_3 . .	*Yellow crystalline* .	precipitate, (Potash chloride, or ammonia-chloride of platinum) sparingly soluble in water and in acids,

		but readily dissolved when heated with an excess of potash or of ammonia. The presence of HCl is almost necessary to this reaction.
KCl or NH_4Cl . .	. . . . .	especially if the liquid contains a little HCl,
	Yellow crystalline .	precipitate (ammonia or potash chloride of platinum)($NH_4Cl, PtCl_2$), but the solution must be pretty concentrated
Hg_2O, NO_5 .	*Orange yellow* precipitate	(See Table IV, Observation *w*)
SnCl .	*Reddish-brown* precipitate,	resembling Fe_2O_3. The superincumbent liquid remains reddish for some time, in more concentrated solutions, a dark brown coloured liquid is produced, but no precipitate. (See Table IV., Observation *x*)

The most characteristic tests for platinum are KO; NH_3—KCl or NH_4Cl—HS.

POLYSULPHURETS. (*See* HYDROSULPHURIC ACID.)

POTASSIUM.

Potassium is a metal which, when freshly cut, has the appearance of lead. At the ordinary temperature it is as soft as wax, but at 32° Fahr. it is brittle and crystalline. By exposure it becomes immediately tarnished and oxydised, wherefore, it is always kept immersed in mineral or Persian naphtha. Its specific gravity is 0·865, it is, therefore, considerably lighter than water. Except mercury, it is the most fusible of all metals; it melts completely at 150° Fahr., and distils over at a low red heat. Potassium has a very great affinity for oxygen, by which it is converted into potash; but it is not altered by exposure in perfectly dry oxygen. It

decomposes water at the ordinary temperature, so that when projected on water, it absorbs the oxygen of it with such violence as to become incandescent, and inflames the hydrogen of the water, which then contains potash. Thrown upon mercury, provided the latter be quite dry, it amalgamates quietly with it, but in the open air it runs on the surface of the mercury, on account of the moisture of the atmosphere and of the hydrogen evolved.

OXYDE AND SALTS OF POTASH.

(KO)

Pure potash (potasse à l'alcohol) is ordinarily in fused cakes of a white colour, extremely caustic, alkaline, and deliquescent Caustic potash (potasse à la chaux) is ordinarily cast into cylindrical sticks.

Potash feels unctuous to the touch, because it dissolves the epidermis of the skin. It deliquesces in the air, combines with water, and the combination is attended with an evolution of heat. The salts of potash are soluble in water, and so are likewise (more or less) all the precipitates which reagents produce in their solutions: hence the necessity of operating upon concentrated liquors. Potash salts are not precipitated by alkaline carbonates, nor by sulphurets, nor by ferrocyanides

TESTS AND REACTIONS

$\bar{T},2HO$	*White* .	produces, in a neutral solution, a crystalline precipitate ($KO,HO,\bar{T}$), which, in concentrated solutions, and provided $\bar{T},2HO$ is in excess, separates immediately after violent shaking If the solution operated upon is acid, it must be neutralised with pure NaO,CO_2; after which an excess of $\bar{T},2HO$ is added. (See Table VI., Observation *r, s, t.*)
$PtCl_2$ . .	*Yellow* . .	precipitate ($KCl+PtCl_2$), in concentrated solutions. The formation of

		this precipitate is promoted by an addition of alcohol, and by a few drops of HCl (See Table VI, Observation *u*, Table XX., Observation *d*)
$SiFl_2,HFl$ . .	*White* . .	gelatinous precipitate, almost transparent.

Blowpipe.—Heated before the blowpipe, or mixed with alcohol, and inflamed, a weak *purple flame* is produced, which is completely obliterated by the slightest trace of soda, in which case the flame is yellow, (See Table I., B, Observation *d.*) Potash salts are not precipitated by alkaline carbonates, nor by sulphurets, nor by ferrocyanides.

The most characteristic reactions for potash are $PtCl_2$ + alcohol—$\overline{T}$,2HO,—the colour of the flame.

PROTOXYDES. (*See* the Metal or the Metalloid.)

RACEMIC ACID (PARATARTARIC ACID), AND RACEMATES.

($C_4H_2O_5$, 2HO, or $\overline{R}$, 2HO.)

This acid is also known under the name of *paratartaric acid,* and is a remarkable instance of an isomeric form of tartaric acid, these two acids have the same formula.

Racemic acid is solid, and crystallises more easily than tartaric acid. Its crystals, like those of tartaric acid, have the form of oblique rhombic prisms, and contain 2 equivalents of water, one of which may be eliminated by exposure to a heat of 212° Fahr. These crystals effloresce in the air.

Racemates interfere with the precipitation of alumina, of peroxyde of iron, and of protoxyde of manganese, by alkalies, precisely as is the case with the tartrates.

TESTS AND REACTIONS.

CaCl . . .	*White* . .	precipitate ($CaO,\overline{R}$), almost insoluble in solution of sal-ammoniac

		but immediately soluble in a concentrated solution of caustic KO, reprecipitated by boiling, and redissolved by cooling This reaction is a distinctive character, of the presence of both tartaric and racemic acids, either jointly or separately.
$CaO + aq$		(Lime-water), poured in excess in the solution of racemic acid, or of a racemate, so as to impart an alkaline reaction, produces a
	White . . .	precipitate (CaO,R)), soluble in HCl, and reprecipitated therefrom by NH_3
CaO,SO	*White* .	precipitate (CaO,R̄) This precipitate is not always produced immediately. but in the course of about a quarter of an hour, the liquor becomes turbid, and subsequently deposits a white precipitate This reaction distinguishes *racemic acid* and *racemates* from *tartaric acid,* and *tartrates, which are not precipitated at all by* CaO,SO_3

The principal tests are CaO,SO_3, which produces a white precipitate in solution of racemic acid and of racemates, but none in those of tartaric acid and tartrates.

The white precipitate produced by *salts of lime*, and which is insoluble in sal-ammoniac, is also very characteristic, since the white precipitate produced in the same circumstances in solution of tartrates is soluble in sal-ammoniac.

RHODIUM.

(Rh)

Rhodium is a white metal, or a gray powder, much less ductile than platinum, very hard and nearly as infusible as iridium; for it scarcely softens before the flame of the oxy-

hydrogen blowpipe. Its specific gravity is 10·64. It is not altered by exposure, except it be heated, in which case it becomes oxydised. The same result is obtained when it is fused with nitrate, or with bisulphate of potash. When pure it is not attacked by any of the ordinary acids, *not even by aqua regia,* but when alloyed with other metals the latter acid dissolves it rapidly.

TESTS AND REACTIONS FOR THE SALTS OF SESQUIOXYDE OF RHODIUM.

KO . . .	*Yellowish-brown*	precipitate (Rh_2O_3,HO), which deposits on boiling
NH_3 . .	*Yellow* .	precipitate, after some time.
KO,CO_2 or NH_4O,CO_2	*Yellow* . .	precipitate, after some time
SnCl	*Red* . . .	colour
KI . .	*Red* . . .	colour
HS . .	*Brown* . .	precipitate, slowly.
NH_4S .	*Brown*	precipitate, insoluble in excess.
Bar of Metallic zinc	*Black* . .	precipitate of metallic rhodium

Hydrogen gas in the cold reduces the solutions of rhodium.

SELENIUM.

(Se.)

Selenium is a rather rare substance, of a reddish-brown colour, and of an imperfectly metallic appearance. Its fracture is vitreous and conchoidal, has a leaden hue, and is somewhat translucent, it is brittle, and when reduced to fine powder it is red, but by boiling, the powder agglomerates and becomes black. (See Table I.—A, Observation *v.*) Its specific gravity is 4·3. It is insoluble in *water,* but soluble in *concentrated* SO_3,HO, and the solution has a green colour; it may be

reprecipitated therefrom by water. Selenium melts at the temperature of boiling water, and at 650° Fahr. it boils and sublimes in yellow fumes resembling S, these fumes condense in the form of a red powder, which very much resembles vermilion. This powder is very inflammable, the result of its combustion being selenious acid (SeO_2), and this combustion is accompanied with a most fœtid odour of rotten cabbage. It dissolves in *nitric acid* and in *aqua regia*, and is thereby converted into selenious acid (SeO_2). *Hydrochloric acid* does not attack it. It dissolves also in solution of KO, but less readily than sulphur. Fused with KO, NO_5, it is converted into seleniate of potash. It forms with oxygen three compounds, namely, oxyde of selenium (SeO), selenious acid (SeO_2), and selenic acid (SeO_3).

The properties of the first of these oxydes are scarcely known; it is a gas which is produced when selenium is heated in oxygen, but without inflaming it. It has the characteristic odour of decayed radish.

SELENIOUS ACID AND SELENITES.

(SeO_2)

Selenious acid, is a solid, very soluble substance, which crystallises in long needles, or large prisms, which volatilise into a yellowish gas without undergoing previous fusion when heated. The principal tests for selenious acid and selenites are the following:

BaCl . . .	*White* . .	precipitate (BaO,SeO_2), soluble in acids.
HS	*Lemon yellow*	precipitate (Se,S) in acidified solutions, it becomes
	Red . . .	by drying.
NH_4S . . .	*Yellow* . .	precipitate, very soluble in an excess of the reagent.
SO_2 or NaO, SO_2 .	*Red* . . .	powder, which is reduced selenium.

A bar of zinc immersed in slightly acid solutions, precipitates selenium in the form of a red powder.

Before the blowpipe selenites behave like seleniates.

SELENIC ACID AND SELENIATES.

(SeO_3.)

Selenic acid is a liquid which has a great affinity for water; it can dissolve a great number of metals, and even gold, but not platinum, and if mixed with HCl it forms a kind of *aqua regia*, and disengages chlorine.

TESTS AND REACTIONS.

BaCl .	*White* . .	precipitate, insoluble in acids, except HCl be used, and heat applied, in which case the seleniate of baryta is decomposed, and chlorine evolved.
HS . . .	*Nothing.*	
NaO,SO_2	*Nothing* . .	in the cold (See Table XXXII, Observation *b*)
Mixed with sal-ammoniac, and heated .	*Red powder*	which is a sublimate of selenium. (See Table I, C, Observations *d* and *e*)

Heated with NaO,CO_2, on charcoal before the blowpipe, an odour of decayed radish is evolved.

SESQUIOXYDE OF IRON (*See* IRON [PEROXYDE]).

SILICA—SILICATES—SILICIC ACID.

(SiO_3.)

Pure silicic acid is a white, gritty powder, which, as found in nature, is insoluble in water. Silicic acid, however, exists in two isomeric states, the one, which is insoluble in water and in acids, *except hydrofluoric acid;* the other, which is

soluble both in water and in acids By evaporating to perfect dryness the soluble form of silicic acid, it is converted into its insoluble modification. When a mixture of insoluble SiO_3, and an excess of KO,CO_2 or NaO,CO_2, is fused, an effervescence is observed, which is due to a disengagement of CO_2 from the alkaline carbonate employed, and the result is a basic silicate of alkali, soluble in water. If the mass or the liquor be then treated by an acid, SiO_3 is separated in a gelatinous form, in which state it immediately dissolves when boiled with solution of caustic or of carbonated potash. (See Table XXI., Observation *a*) But if the gelatinous silicic acid separated by an acid be evaporated to perfect dryness, it becomes completely insoluble in water and in acids, though it remains soluble in solutions of potash and of carbonate of potash , if, however, it be not only dried, *but ignited,* it becomes insoluble, or only triflingly soluble even in these menstrua.

The silicates of the alkalies alone are soluble in water. (See Table XXIX., Observation *c*.)

TESTS AND REACTIONS.

The silicates which are insoluble in water and in acids, must be fused with three or four times their weight of KO,CO_2, or NaO,CO_2, or better still with a mixture of both, and the fused mass being decomposed by HCl, is then evaporated *to perfect dryness,* in order to convert the gelatinous SiO_3 into the insoluble modification, the dry mass being then digested for a short time with strong HCl, and then treated by boiling water, leaves a perfectly white and gritty powder, insoluble in water and in acids, and which is SiO_3. (See Table I., C, Observation *c*; Table X., Observation *e*; Table XXI., Observation *a*, Table XXIX., Observation *c*.)

HFl dissolves SiO_3, in whatever state it may be.

All the silicates even by exposure to a red heat resist decomposition.

Blowpipe.—Mixed with NaO,CO_2, and fused on a hook of platinum before the blowpipe, pure silicic acid produces a *clear bead.* It is absolutely necessary not to use with the SiO_3 more than at most its own bulk of NaO,CO_2, since if too much or too little NaO,CO_2 be employed, the bead is opaque after cooling. It is evident that if the silicic acid is not *pure*, it will impart to the bead obtained the characteristic colour of the oxyde by which it may be contaminated. (See Table I., B, Observation *h*; Table XI., Observation *j*.)

The following natural silicates are completely decomposed by HCl. All the minerals marked with an asterisk (*) in the following list, are decomposed with difficulty by HCl, and only when reduced into fine powder, a concentrated acid must then be used with the help of heat. All the others at once dissolve into a gelatinous mass when treated by HCl.

	Combination of
Amphigène. (*See* Leucite)	
Apophyllite . .	potash, silica, lime, and water.
*Anorthite	silica, alumina, lime, magnesia
Analcime . .	silica, soda, water.
Alumocalcite . .	silica, alumina, lime, water.
Allophane . .	alumina, silica, water.
*Allanite (*See* Cerine)	
Brewsterite . .	silica, alumina, strontia, baryta, lime, and water.
Botryolite[1] . .	silica, boracic acid, lime, and water
Copper Malachite. (*See* Malachite)	
Cronstedtite . .	silica, oxyde of iron, water (hydrated silicate of iron).

[1] It differs from datholite only because it contains one equivalent of water more than the latter.

	Combination of
Cancrinite[1] . . .	soda, silica, and lime.
Chronikrite . . .	silica, magnesia, alumina, lime, protoxyde of iron, and water.
Chabasite . . .	silica, alumina, lime, water, with a little potash.
*Cerite	hydrated silicate of peroxyde of cerium.
*Cerine . .	silicate of alumina and of cerium, of iron and of lime.
Davyne	silica, alumina, lime, iron, water.
Dysclasite . . .	silica, lime, water.
Datholite[2] . . .	silica, boracic acid, lime, water.
Dioptase	silicate of copper, water.
Eleolite . . .	silica, alumina, lime, potash, soda, water.
Eudialite[3] . . .	silica, soda, zirconia, lime, oxyde of iron, manganese, hydrochloric acid, and water.

[1] Cancrinite is thought by some mineralogists to be the same mineral as sodalite. The analysis of cancrinite and sodalite gives the following results.

	Sodalite				Cancrinite.
	From Greenland		From Vesuvius		Mines in Siberia
Soda	25 00	25·50	26 55	20 96	24 47
Silica .	36 00	38 52	35 99	50 98	38 40
Alumina .	32 00	27 48	32 59	27 64	32 04
Lime .	0 00	2 70	0 00	0 00	0 32
Hydrochloric acid	6 75	3 06	5·30	1 29	
Protoxyde of iron .	0 25	1 00	0 00	0 00	
Volatile matter	0 00	2 10	0 00	0 00	Loss 4 77
	100 00	100 00	100 00	100 00	100 00
	Eckeberg	Thomson	Arfwedson	Watchmeister and Berzelius	Hoffmann

[2] This mineral differs from botryolite only because it contains one equivalent less water.

[3] According to Beudant, double silicate of zirconia and of soda, lime and iron.

	Combination of
Electric Calamine [1].	. silica, oxyde of zinc, and water.
*Epistilbite [2] . .	. silicate of alumina and of lime; water.
Gehlenite . .	. silica, alumina, lime, oxyde of iron.
Gadolinite . .	. yttria, glucina, silica, oxydes of cerium and iron.
Helvine . . .	. silica, glucina, alumina, protoxyde of iron and of manganese.
Hauyne . . .	. potash or soda, silica, alumina, lime, and sulphuric acid.
Harmotome. (*See* Potash Harmotome)	
Hisingerite . .	. silicate of protoxyde and of sesquioxyde of iron + 6 equiv of water
*Heulandite . .	. tersilicate of alumina and of lime
Ilvaite .	. silica, protoxyde of iron and lime
Laumonite . .	. silica, alumina, lime, and water.
Leucite . . .	. potash, silica, and alumina.
Lazulite . .	. silica, alumina, lime, oxyde of iron, magnesia, soda, and sulphuric acid.
Marceline .	anhydrous silicate of binoxyde of manganese.
Mesolite .	. silica, alumina, lime, soda, and water
Mesole . . .	. silica, alumina, lime, soda, and water
Mesotype . .	. silica, soda, and water.
Melilite . . .	. silica, magnesia, lime, oxyde of iron

[1] This mineral must not be confounded with ordinary calamine, which is a carbonate of zinc, whilst electric calamine is a hydrated silicate of zinc.

[2] The symbol is $(4Al_2O_3,3SiO_3) + (CaO,3SiO_3) + 6HO$; That is to say: 4 equiv of tersilicate of alumina, 1 equiv of tersilicate of lime, and 6 equiv. of water, being the mean of two analyses by Walmsted and Thomson. Epistilbite is dissolved by concentrated HCl, except a small gritty residue of silica.

	Combination of
Meerschaum . .	. magnesia, carbonic acid, silica, water, a little alumina, and traces of manganese and of lime
Meionite . .	. silica, alumina, lime, and soda (?)
Malachite (copper) .	. hydrated carbonate and silicate of copper
Manganese . .	. ferruginous silicate, or Troostite
Manganese .	(silicate of protoxyde of)
Manganese .	hydrosilicate. (*See* Opsimose.)
Natrolite . .	. silica, soda, and water
Nepheline . .	. soda, silica, and alumina
Nosian . . .	. sesquisilicate of alumina and soda
Okenite . . .	. silica. lime, soda, and potash, oxyde of iron, oxyde of manganese, water.
Opsimose . .	. hydrosilicate of manganese
Orthite . .	. silica, alumina, oxydes of iron, cerium, lanthanum, manganese, lime, yttria, magnesia, and a small quantity of water.
Potash Harmotome	. silica, alumina, baryta, potash, and water (sometimes lime)
Pectolite . .	. silica, lime soda, potash, water and oxyde of iron
*Pyrosmalite . .	. lime, tersilicate of oxyde of iron and of manganese.
*Pitchblende . .	. uranium ore, containing about three per cent of silica, probably in a state of mechanical, not of chemical, combination.
Scolezite . . .	. silica, alumina, lime, soda, and water.
Sodalite . . .	. soda, silica, and alumina, with a small quantity of hydrochloric acid.
Spar (tabular) . .	. silica and lime.

	Combination of
Sideroschisolite . .	protoxyde of iron, silica, alumina, and water.[1]
*Stilbite [2] . .	tersilicate of alumina + silicate of lime + 6 equivalents of water
*Sphene . . .	tersilicate of lime, and titaniate of lime.
Scapolite. (*See* Meionite.)	
*Titanite (*See* Sphene.)	
Troostite . .	protosilicate of manganese, sesquisilicate of iron, and ferruginous silicate of manganese.
Tabular Spar. (*See* Spar.)	
Wernerite . .	silica, alumina, lime, soda
Yenite	silica, peroxyde of iron, protoxyde of iron, lime.

The following natural silicates altogether resist the action of HCl, even after fine pulverisation, and are decomposed only by fusion with about four times their weight of NaO,CO_2, or KO,CO_2.

	Combination of
Albite . . .	silica, alumina, and soda
Achmite . . .	soda, silica, and peroxyde of iron.
Amphibole (Hornblende) .	silica, magnesia, and lime (The silica is partly replaced by alumina.)
Anthophyllite . .	bisilicate of magnesia, and bisilicate of iron.
Axinite . . .	silica, alumina, lime, oxyde of iron, and manganese.
Barytic Harmotome . .	silica, alumina, baryta, water
Beryl. (*See* Emerald.)	

[1] This mineral is probably a basic silicate of iron, the alumina being, probably, an accidental constituent.

[2] The difference between stilbite and epistilbite consists in the latter containing one equivalent of water less than the former

	Combination of
Brucite. (*See* Condrodite)	
Condrodite . . .	silica, magnesia, peroxyde of iron, and generally fluorine.
Carpholite . .	silica, alumina, protoxyde of manganese, protoxyde of iron, lime, fluorine, and water. (The lime and fluorine are looked upon as accidental.)
Chlorite (Venetian Talc)	silica, magnesia, oxyde of iron, alumina, (potash ?), water.
Dichroite	silica, alumina, magnesia, oxyde of iron, and manganese
Diallage . .	hydrated bisilicate of magnesia.
Epidote .	silica, alumina, protoxyde of iron, and lime
Emerald . .	glucina, silica, and alumina.
Euclase . . .	glucina, silica, alumina, oxyde of iron, and tin. (The two latter are accidental constituents.)
Felspar .	silica, alumina, potash, and soda.
Garnet . .	silicate of protoxyde of iron, and silicate of alumina.
Harmotome. (*See* Barytic Harmotome)	
Hornblende (*See* Amphibole)	
Idocrase . .	silica, lime, alumina. (Protoxyde of iron, of manganese, and magnesia are accidental.)
Iolite. (*See* Dichroite)	
Labradorite . .	silica, alumina, lime, soda, oxyde of iron.
Lepidolite. (*See* Mica, common.)	
Mica (Magnesia) . .	silica, alumina, potash, magnesia, peroxyde of iron.
Mica (common, or Lepidolite) . .	silica, alumina, potash, oxyde of manganese, lithia.
Manganese Spar [1] . .	bisilicate of manganese.

[1] The bisilicate of manganese is partly dissolved by HCl.

	Combination of
Oligoclase. (*See* Soda-Spodumene.)	
Olivine	anhydrous silicate of magnesia. (Protoxyde of iron accidental ?)
Obsidian . .	soda, potash, silica, alumina, oxyde of iron.
Pitchstone . .	soda, potash, silica, alumina, oxyde of iron.
Pumicestone . .	soda, potash, silica, alumina, oxyde of iron.
Petalite .	silica, alumina, lithia.
Prehnite . .	silica, alumina, lime, and water
Pinite . .	potash, silica, alumina, and soda; magnesia, and oxydes of iron and of manganese.
Pyroxene . . .	silica, lime, protoxyde of iron, and sometimes manganese and alumina.
Phenakite .	silica, glucina, alumina, and magnesia. (Essentially bisilicate of glucina)
Picrosmine .	silica, magnesia, water (protoxyde of iron, and of manganese, accidental). Essentially hydrated bisilicate of magnesia.
Ryacolite .	silica, alumina, potash, soda.
Soap-stone (*See* Steatite.)	
Spodumene (Soda) .	bisilicate of alumina and tersilicate of lime, magnesia, potash, and soda.
Spar (Chatoyant) (*See* Diallage)	
Steatite (Soap-stone)	silica, alumina, magnesia, and water.
Serpentine . .	hydrated sesquisilicate of magnesia.
Talc. (*See* Chlorite.)	
Tordawalite . .	silica, alumina, peroxyde of iron, magnesia, phosphoric acid, and water. (Partly decomposed by HCl.)
Tourmaline .	silica, alumina, oxyde of iron and lime with small portions of magnesia, potash soda, and boracic acid.
Topaz	alumina, silica, and fluorine.

The following natural silicates cannot be decomposed by HCl, nor even by fusion with NaO,CO_2, or KO,CO_2, but are decomposed by fusion with pure KO.

	Combination of
Andalusite . .	essentially silicate of alumina
Cymophane (Chrysoberyl)[1]	alumina, glucina, silica, lime.
Cyanite	silicate of alumina.
Staurolite	alumina, silica, and oxyde of iron
Zircon	silicate of zirconia

The following natural silicates, heated by themselves or with soda, borax, or microcosmic salt, behave as follows:

	Alone	*With Soda*	*Borax*	*Microcosm*
Achmite.	Shiny black bead	Reddish-brown	Reddish-brown	
Albite .	Transparent	Transparent		
Allanite .	Black shiny bead on charcoal, with effervescence		Blood red in oxydising flame when hot, yellow when cold	
Allophane	Swells up, and falls to powder, but does not fuse (the flame is tinged with green)	Colourless bead	Colourless bead	
Alumocalcite	Becomes grey and opaque	Colourless glass	Colourless bead.	Skeleton of silica left
Amphibole	With effervescence, black glass	Colourless	Colourless	Opaline.
Amphigene	Infusible	Transparent, vesicular (effervescence)	Transparent.	
Analcime .	White and opaque (on charcoal), and then transparent	Transparent	Difficult of fusion	

[1] Silica and lime are not essential constituents.

	Alone.	*With Soda*	*Borax*	*Microcosm*
Andalusite .	Infusible.	Almost infusible	Almost infusible	
Anorthite	Almost infusible			Opaque glass
Anthophyllite	Infusible	Reddish-brown	Reddish-brown	Skeleton of SiO_3
Apophyllite	Swells and fuses into a vesicular glass	Transparent	Transparent	
Axinite	Dark green glass (intumescence)	Almost infusible (reddish, or black, glass)	Almost infusible (reddish, or black, glass)	
Beryl	Looks like mother of-pearl	Colourless	Colourless	Opaline
Botryolite	White glass	White, transparent	Transparent	
Brewsterite	Fuses with effervescence	Almost transparent, colourless	Almost transparent	Skeleton of SiO_3
Calamine (electric)	Decrepitates instantly, but does not fuse	Colourless	Colourless	
Cancrinite	Colourless glass (intumescence)	Colourless glass.	Colourless glass, but with difficulty	
Carpholite	Brown opaque glass (intumescence)	Transparent	In the outer flame amethyst, in the inner flame green when cold, colourless	
Cerine (*See* Allanite)				
Cerite	Infusible (decrepitates)	Almost colourless glass	Orange glass, when hot, almost colourless when cold.	
Chabasite	White opaque	Transparent	Transparent	
Chlorite	Difficult to fuse	Dingy glass		
Chromicrite .	Grey glass (effervescence)	Brownish transparent glass	Reddish glass, when hot.	
Condrodite .	Infusible, but decolorised	Transparent yellowish glass	Transparent Reddish-brown glass, when hot, yellowish, when cold	
Cronstedtite	Infusible	Black opaque glass.	Black opaque mass	

	Alone	*With Soda*	*Borax*	*Microcosm*
Datholite	Transparent pink bead (intumescence).			
Davyne .	White opaque bead	Colourless glass	Colourless glass	Opaline glass
Diallage	Infusible	Reddish-brown glass.	Almost infusible.	Skeleton of silica
Dichroite	Almost infusible	Colourless.	Colourless	
Dioptase . .	Infusible, but becomes black in the outer, and red in the inner, flame	Greenish glass.	Greenish bead	
Dysclasite	Almost infusible	Semitransparent glass, with intumescence	Colourless bead	Colourless bead
Eleolite . .	White opaque glass	White glass	White glass	
Emerald .	Clouded	Colourless	Transparent greenish glass	Transparent greenish glass
Epistilbite	White vesicular enamel	Colourless		
Euclase .	Almost infusible, except at the edges	Colourless	Colourless glass (slowly)	
Eudialite	Green scories			
Felspar	Almost infusible.	Colourless	Colourless	
Gadolinite		Dingy-green glass	In reducing flame, bottle-green glass	
Garnet	Black glass	Dark-grey mass	Dark yellowish glass (slowly)	
Gehlenite	Infusible	Yellowish dingy glass	Reddish-brown glass, when hot, yellowish, when cold	
Harmotome (Barytic) .	Colourless glass	Colourless glass	Colourless glass	
Haüyne	Opaque mass	Yellow glass.	Colourless glass, when hot, yellow, when cold	Opaline yellow glass Skeleton of SiO_3.
Helvine .	Yellow opaque mass (on charcoal)		Yellow glass, when cold	

	Alone	*With Soda.*	*Borax*	*Microcosm*
Heulandite	Intumesces, and fuses with phosphorescent light.			
Hisingerite	Opaque black globule		Yellowish glass	
Hornblende	(*See* Amphibole)			
Idocrase .	Yellowish transparent glass	Greenish mass	Greenish glass	
Iolite	Almost infusible, except at the edges	Colourless	Colourless glass	
Labradorite	Colourless	Colourless	Colourless	
Laumonite	White opaque spumous mass	Colourless	Colourless	
Lazulite	White glass	Opaque greenish grey glass, reddens on cooling		
Lepidolite .	Infusible, except lithia is present	Colourless		
Leucite .	Infusible	Colourless (vesicular)	Colourless	
Lievrite . .	(*See* Yenite)			
Malachite (Copper)	Black scorios	Green glass	Green (bead of copper)	
Manganese (Hydrosilicate) . .	(*See* Opsimose)			
Manganese (Ferruginous silicate) .	(*See* Troostite)			
Manganese (Spar)	Reddish-brown, or black globule	Turquoy mass.	Violet glass.	
Marceline .	Fuses, greyish black globule	Turquoy mass Violet glass.		
Meïonite .	(*See* Scapolite)			
Mellilite .	Greenish glass.			
Mesotype .	Opaque glass.	Colourless.	Colourless	
Mica (Magnesian) .	Fuses into a white opaque mass.			

	Alone	*With Soda*	*Borax*	*Microcosm.*
Mica (common)	(*See* Lepidolite.)			
Natrolite .	(*See* Mesotype.)			
Nepheline	Colourless vesicular glass.	Colourless	Colourless	Dissolves, leaves skeleton of SiO_3
Okenite	Almost infusible	Semi-transparent glass	Colourless	Colourless
Oligoclase (or Soda Spodumen). .	Fuses with difficulty.			
Olivine	Infusible	Dingy glass	Yellowish glass	
Opsimose	Green glass, in the inner flame, black glass, in the outer flame	Green glass	Amethyst	
Orthite .	Black blistering glass	Turquoy mass	Red bead, hot, yellow bead, cold	
Pectolite	White glass	Colourless	Colourless	
Petalite	Difficultly and only on the edges	Colourless	Colourless	
Phenakite	Difficultly	Colourless		
Picrosmine	Infusible	Greenish mass		
Pinite	Infusible, or fuses into a black glass	Colourless	Colourless.	
Pitchblende	Infusible		Deep yellow glass	
Pyrosmalite	Black glass		Brownish, when red hot, yellowish, when cold	
Quartz . .	Infusible	Colourless glass		
Scapolite.	Fuses with intumescence	Colourless.	Colourless.	
Serpentine	Almost infusible		Almost white glass.	
Sideroschizolite	(*See* Cronstedtite)			
Sodalite . .	Infusible: other varieties fuse into a colourless globule.	Colourless	Colourless	

	Alone	*With Soda*	*Borax*	*Microcosm*
Sordawalite	Fuses difficultly into a dark globule	Blackish slag	Green glass	
Spar (Tabular)	Fuses, but with difficulty	Colourless glass	Colourless glass	
Spar (Chatoyant)	(*See* Diallage)			
Sphene	Almost infusible		Yellowish-green glass	Fuses slowly into opaline glass
Spodumen	Fuses into a transparent glass	Colourless glass	Colourless glass	
Steatite	Infusible, but turns black			
Stilbite	Colourless blistering glass	Colourless glass	Colourless glass	
Talc .	Difficultly fusible	Dingy glass	Reddish-brown, when hot, yellowish, when cold	
Titanite	(*See* Sphene)			
Topaz	Infusible	Colourless.	Colourless.	
Tourmaline	Black scorios	Colourless	Colourless	
Troostite	Fuses on the edges	Turquoy blue	Amethyst	
Wernerite	(*See* Scapolite)			
Yenite	Black glass.		Dark-green opaque glass	
Zeolites		Colourless glass		

SILVER.

(Ag.)

Silver is the whitest of all metals, and has a beautiful metallic lustre. It is very malleable and ductile; a little harder and more fusible than gold; its specific gravity is 10·5. It melts at a bright red heat (about 1873° Fahr.), and then absorbs oxygen from the air, but it parts with it again on cooling. It has no odour, no flavour, and it does not oxydise by exposure.

Dilute *sulphuric acid* does not attack silver, but if the acid be concentrated then the silver is dissolved, and sulphurous acid is disengaged.

Phosphoric acid, in the wet way, has no action on silver, but in the dry way silver is attacked by that acid.

Hot concentrated HCl has an action on silver, especially if it be put in contact with platinum at the same time ; the result is a basic chloride of silver, hydrogen being disengaged at the same time.

A blade of silver plunged into HS in the state of gas, or of solution, immediately becomes black.

The *vegetable acids* have no action on silver.

The best solvent of this metal is NO_5, the result being nitrate of silver, and pure nitric oxyde is evolved.

Silver is not altered by the *caustic alkalies,* nor by the *carbonates, nitrates,* and *chlorates of alkalies,* wherefore silver crucibles are often used for the fusion of substances with these compounds; yet common salt, kept fused for some time in silver, always produces a certain quantity of chloride of silver.

OXYDE AND SALTS OF SILVER.

(AgO)

Protoxyde of silver is a brown powder, which turns black by exposure; it is a powerful base, which completely saturates the strongest acids, wherefore nitrate of silver is neutral to test papers ; *nitrate of silver* reddens litmus paper only when it contains *free nitric acid.* There is no hydrate of that oxyde, though the yellowish tinge which it assumes at the moment of its being precipitated, seems to indicate that such a hydrate exists—at any rate, if there be really such an oxyde, it is rapidly decomposed whilst washing it. It is slightly soluble in pure HO.

The salts of silver are colourless if their constituent acid is colourless. They have a metallic and astringent taste;

they are blackened by exposure to daylight. Several salts of silver, but not all, have no reaction on test papers. Most of the salts of silver are decomposed by a red heat.

TESTS AND REACTIONS.

HS or . . NH_4S .	*Black* . .	precipitate of sulphuret of silver in neutral, in acid, and in ammoniacal solutions.
NH_3	*Light brown,*	A very small quantity produces a or rather *drab* precipitate in neutral solutions, *immediately soluble in the slightest excess of the reagent.* The presence of ammoniacal salts prevents the formation of that precipitate, and, consequently, if the solution contains a free acid, ammonia will produce no precipitate.
HCl	*White curdy*	and metallic chlorides, produce an abundant precipitate of $AgCl$, insoluble in dilute acids; immediately soluble in aqueous solution of NH_3, reprecipitated by supersaturation with nitric acid (characteristic) (See Table XIV, Observation *b*, Table XVII., Observation *f*.) Exposed to solar light, the precipitate becomes *grey* or *purple*. If only a small quantity of silver is present, *no precipitate* is produced, but the liquor is rendered opaline.
KO,CrO_3 . .	*Crimson red*	precipitate of AgO,CrO_3; soluble in dilute NO_5, in NH_3, and in a large quantity of water.
KO,CO_2 or NaO,CO_2 or NH_4O,CO_2.	. . .	White precipitate (AgO,CO_2); soluble in an excess of NH_4O,CO_2, and of course in NH_3.
KI	*Yellowish white* . .	precipitate (AgI); insoluble in an excess of NH_3.
FeO,SO_3 . .	*White* . .	precipitate of metallic silver.

Metallic silver is reduced from the solution of the salts of silver by Cu—Hg, and by Zn.

The most characteristic tests for silver are the reactions with HCl and NH_3.

SODIUM

(Na)

Sodium is a silver-white metal possessing considerable metallic lustre, and having a great resemblance with potassium. Like the latter metal it is as soft as wax at the ordinary temperature, but below the freezing point it is brittle, it melts at 194° Fahr., and volatilises at a low red heat. Sodium is rapidly oxydised by exposure, and on that account it is usually kept in Persian or mineral naphtha or petroleum. Its specific gravity is ·972, it therefore floats when thrown into water, which it violently decomposes, but without inflammation, except its rapid motion at the surface of the liquid be arrested either by putting but little water, so that it may touch the bottom of the vessel, or by thickening the water with gum or other mucilage, in which case it burns with a characteristic yellow flame.

OXYDE OF SODIUM (SODA) AND SALTS OF SODIUM.

(NaO)

Pure oxyde of sodium or soda (Soude à l'alcohol) is very similar to potash. It is ordinarily met with, like the latter alkali, in fused cakes, which are extremely caustic and alkaline; it also deliquesces by exposure, but instead of remaining syrupy, as is the case with deliquesced potash, it soon solidifies by absorbing the CO_2 of the air. All the salts of soda are soluble in water, and form no precipitate with any of the reagents which act on KO.

TESTS AND REACTIONS.

KO,SbO_5	*White crystalline*	When this reagent has been dissolved in cold water, it produces, in even dilute salts of soda, a precipitate (NaO,SbO_5), soluble in about 300 parts of water, but the operator must make sure that no other oxyde besides KO is present, and the liquor must be slightly alkaline: if acid, the acid should be neutralised with KO to slight alkaline reaction, before testing.
KO,IO_7	*White*	precipitate, sparingly soluble in water, the solution, to be tested, should be very concentrated.

Blowpipe.—Evaporated nearly to dryness, mixed with alcohol, and inflamed, or a portion of the dry mass heated before the blowpipe, produces a characteristic yellow flame. (See Table I., B., Observation *d.*)

STRONTIUM.

(St.)

Strontium is a white metal of a feeble metallic lustre, which is rapidly oxydised by exposure, and is thereby converted into strontia.

OXYDE OF STRONTIUM (STRONTIA) AND SALTS OF STRONTIUM.

(StO.)

Pure oxyde of strontium or strontia has the greatest resemblance with baryta, but it is not so heavy. It combines very energetically with water, with evolution of heat, the

result being hydrate of strontia (StO,HO), which is less soluble in water than baryta, 160 parts of that menstruum being required to dissolve it. Boiling water dissolves a larger proportion of strontia, but it crystallises therefrom on cooling in the form of thin quadrangular tables. Its aqueous solution must be kept in well-stoppered bottles, as it absorbs rapidly the carbonic acid of the air. Its salts are colourless, and most of them are insoluble, or only sparingly soluble in water.

TESTS AND REACTIONS.

SO_3	*White* . .	and the soluble sulphates, produce a precipitate (StO,SO_3), insoluble in acids and in alkalies, but not quite insoluble in water (about 1 in 4000) Hence it is that solution of sulphate of lime never produces an immediate precipitate in those of strontia.
$(2NaO),HO,PhO_5$	*White* .	precipitate ($(2StO),HO,PhO_5$) in neutral and in alkaline solutions.
NaO,CO_2 . . .	*White* . .	and other alkaline carbonates, precipitate (StO,CO_2) (*See* Table XXV., Observation *a.*)
$NH_4O,\bar{O}$.	*White* . .	precipitate ($StO,\bar{O}$), especially if free NH_3 is present (*See* Table VI., Observation *g*; Table X, Observation *d*)
$SiFl_2HFl$ . .		*No precipitate*, and therefore this reagent serves to distinguish StO from BaO.
$KO,2CrO_3$ . .		*No precipitate*, and therefore this reagent distinguishes StO from BaO. (*See* Table VI., Observation *f.*)
KO . . .	*White* . .	precipitate, soluble in great excess of HO.
Alcohol		dissolved in alcohol, or mixed in powder with alcohol, a carmine *red flame* is produced on inflaming it.

The salts of strontia differ from those of baryta because

$SiFl_2HFl$	. *does not precipitate* the salts of strontia, whilst it does those of baryta
$KO,2CrO_3$ . .	*does not precipitate* (or only after a time) the salts of strontia, whilst a yellow precipitate is at once formed in those of baryta.
Alcohol . .	Mixed with alcohol, and inflamed, the alcohol burns with a carmine flame, if strontia is present, whilst the colour of the flame is scarcely affected by those of baryta. Remember, however, that the salts of calcium impart also a red tinge to the flame of alcohol.

The most characteristic tests for strontia are CaO,SO_3,—$SiFl_2,HFl$—and the *red colour* of the alcoholic flame.

SUBOXYDES OR SUBSALTS. (*See* the METAL OR METALLOID.)

SUCCINIC ACID AND SUCCINATES.

($C_4H_2O_3$, HO, or $\bar{S}$, HO)

Succinic acid forms regular colourless and odourless crystals, of a somewhat nauseous taste, which are soluble in 5 parts of cold, and in about 2 parts of boiling water, and which fuse and volatilise without undergoing decomposition when heated, the fumes evolved are very acrid. Succinic acid is very soluble in alcohol, but sparingly so in ether. Several chemists consider succinic acid as a bibasic acid, and write its formula, $C_8H_4O_6$,2HO.

Most succinates are soluble in water, and are decomposed by a red heat, those of potash, of soda, or of the alkaline earths, are thereby converted into carbonates of these bases.

Succinate of ammonia sublimes without undergoing decomposition.

TESTS AND REACTIONS

Fe_2Cl_3	*Reddish brown*	In neutral solutions of succinates of alkalies, light voluminous precipitate ($Fe_2O_3, 2\bar{S}$) immediately soluble in acids and in ammonia. Wherefore the solution to be tested with Fe_2Cl_3 must be perfectly neutral (*See* Table VIII, Observation 9.)
$PbO,\bar{A}$ . .	*White* . .	precipitate of succinate of lead ($PbO, \bar{S}$), soluble in an excess of the reagent, in NO_5, in $\bar{T}$, 2HO, and in $\bar{A}$,HO, though not so readily.
Alcohol . . + NH_3 + . . BaCl . . .	*White* .	precipitate ($BaO,\bar{S}$), because the earthy succinates are insoluble in alcohol.

SULPHUR.

(S)

Sulphur is a simple substance of a lemon-yellow colour, solid and brittle at the ordinary temperature, tasteless, inodorous, but acquiring by friction a characteristic odour. It is insoluble in water, and scarcely soluble in alcohol or ether, but it dissolves in bisulphuret of carbon, oils, and turpentine. Its specific gravity is 1·987. It melts at 226° Fahr., it has then the colour of amber, and is as fluid as water. If the heat be increased it gradually darkens and thickens, so that at from 430° to 480°, it is quite brown, and so thick and tenacious that the vessel in which the operation is carried on can be turned upside down without spilling; if the heat be further increased it becomes gradually thinner and more fluid, and if it be then rapidly cooled, such as by pouring it into water, it remains for a long time soft and flexible, so that it may be drawn into long threads, in this state it feels

somewhat like a mass of caoutchouc. Sulphur boils and sublimes at about 600° Fahr., and its vapour condensing in minute crystals of a yellow colour, constitutes what is generally known under the name of flowers of sulphur.

Sulphur has a great affinity for oxygen, and burns in it when at a temperature of about 300° Fahr. with a beautiful blue flame and a suffocating odour of SO_2, into which it is thereby converted.

Boiled with NO_5 it gradually disappears, and is converted thereby into sulphuric acid, the *complete* conversion of sulphur into sulphuric acid, however, takes a long time, though if fuming NO_5 be employed it is more rapid.

HCl has no action on sulphur.

Aqua regia converts it into sulphuric acid more rapidly than nitric acid alone.

Solution of KO, or of NaO, or of their carbonates, dissolves sulphur rapidly, especially by boiling; the solution contains sulphuret and hyposulphate of potassium or of sodium.

NH_3 does not act on sulphur.

Fused with an alkaline nitrate, it is converted into a sulphate of the alkali.

SULPHURETS (*See* HYDROSULPHURIC ACID).

SULPHOCYANIDES (*See* HYDROSULPHOCYANIC ACID).

SULPHURETTED HYDROGEN (*See* HYDROSULPHURIC ACID.)

SULPHURETS OF ARSENIC.

There are several sulphurets of arsenic, namely AsS_2 (realgar); AsS_3 (orpiment); AsS_5 (penta, sulphuret of arsenic).

The first is red or orange-red, the two others are yellow. They are not quite insoluble in water, or rather they yield to it some AsO_3 and some sulphuret; 150 parts of water take up, at a boiling temperature, 1 part of orpiment, and 1·5 part of realgar. The sulphurets of arsenic (AsS_3,—AsS_5) obtained by precipitation with HS, are still more soluble in that menstruum. AsO_3 is always contained in the sulphurets of arsenic prepared in that dry way.

The sulphurets of arsenic readily yield unquestionable evidence of the presence of arsenic when submitted to the tests described under the head of AsO_3, namely by treatment in a reduction tube with NaO,CO_2+KCy; or with Marsh's apparatus, or by Reinch's process, or they may be dissolved in NO_5 and their solution thus obtained treated with AgO,NO_5 as mentioned before under the title Arsenious Acid. (See also Table XXVIII., A, Observations *b*, *c*, *d*, *e*, *f*.)

SULPHURIC ACID.—Sulphates.

(SO_3)

Pure anhydrous sulphuric acid is solid at the ordinary temperature, and is in white silky masses resembling asbestus, which evolve thick white fumes when exposed to the air. Ordinary concentrated or monohydrated sulphuric acid, also called oil of vitriol (SO_3,HO), is a heavy, colourless, odourless, oily looking fluid having a great affinity for water, and a most corrosive action upon organic substances, which are thus rapidly destroyed, and charred by it. By mixing it with water a great elevation of temperature is produced.

All the acid and most of the neutral sulphates are soluble in water. The principal exceptions are sulphate of baryta which is quite insoluble, sulphate of strontia and sulphate of lead which are very nearly so, and sulphate of lime which is soluble in about 400 parts of water.

Almost all the basic sulphates are insoluble in water, but dissolve in the dilute acids.

All sulphates are insoluble in alcohol.

All sulphates are decomposed by ignition, and disengage either sulphuric, or sulphurous acid, and oxygen, but the sulphates of alkalies, and of the alkaline earths, and of lead, resist the action of a high temperature. (See Table I., A, Observation *p*.)

TESTS AND REACTIONS.

$BaCl$ or BaO,NO_5	*White*	precipitate (BaO,SO_3) immediately produced, which precipitate is perfectly insoluble in water and in acids. Remember that several barytic salts, which are very soluble in water, are but sparingly so in acids; thus, for example, if the solution under examination contains a large excess of HCl, the addition of $BaCl$ or of BaO,NO_5 will produce a precipitate which may be erroneously taken for one of BaO,SO_3, but which is really nothing else than $BaCl$, which being only sparingly soluble in the acid of the solution, disappears entirely on diluting it with a larger quantity of water. (See Table VII., Observation *s*, Table XXI., Observation *g*.)
$PbO,\bar{A}$	*White*	precipitate (PbO,SO_3), almost insoluble in water, sparingly soluble in dilute NO_5, but completely soluble in concentrated HCl, especially with the help of heat.

Fusion with NaO,CO_2 converts sulphates into carbonates, sulphate of soda being formed thus:

$$CaO,SO_3+NaO,CO_2= \\ CaO,CO_2+NaO,SO_3.$$

Sulphate of lead however, when so treated, is converted into oxyde of lead (PbO), and not into carbonate of lead.

Before the blowpipe sulphates are most readily detected by fusing them with NaO,CO_2 on a charcoal support. If the mass be then removed from the charcoal support, placed upon a clean piece of silver, and moistened thereon with HCl, a brown stain will be produced on the silver, an odour of HS being generally evolved. If the mass is absorbed by the charcoal during the experiment, it is a sulphate of alkali. (See Table I., A, Observation *p.*)

Hyposulphates, sulphites, hyposulphites, and sulphurets behave like sulphates when fused upon charcoal with NaO, CO_2, but treatment with SO_3,HO will have distinguished these compounds from each other, as shown in Table I., C. If the sulphate contained in the combination is one which gives no coloured glass with the fluxes, the fused mass will have a yellow or brownish colour due to the fused sulphuret of sodium which is formed.

Sulphates thrown into fused nitre evolve ruddy fumes of nitrous acid.

To ascertain whether sulphuric acid exists in a solution in a free or uncombined state, Runge's test is excellent, it is as follows: Dissolve 1 part of sugar in about 30 parts of water, moisten a china saucer with the saccharine solution, put a drop of the suspected liquor on the saucer, and heat it by placing the saucer upon a steam bath. If free SO_3 is present, a greenish, or an intensely black spot will be produced, the intensity of which is proportionate to the quantity of free sulphuric acid present.

The most characteristic tests of the presence of sulphates and of sulphuric acid, are BaCl or BaO,NO_5—$PbO,\bar{A}$—and fusion before the blowpipe with NaO,CO_2 on a charcoal support.

SULPHUROUS ACID.—Sulphites.

(SO_2)

At the ordinary temperature and pressure sulphurous acid is a permanent gas of a suffocating odour (that of burning

sulphur), and of a peculiar acid taste. Exposure to the air converts it gradually into sulphuric acid. It is very soluble in water, and more largely still in alcohol, from both of which menstrua it may be expelled by ebullition, but in doing so a portion of it is always converted into SO_3. Neutral sulphites are inodorous. Sulphurous acid is displaced from its combinations by all acids except by carbonic and hydrocyanic acids.

Acid and basic sulphites taste and smell of sulphurous acid The acid sulphites and the sulphites of alkalies are soluble in water, most other sulphites are insoluble therein.

TESTS AND REACTIONS.

HS	*White* . .	precipitate of sulphur (in *acid* solutions)
HCl . . .	*Odour of SO_2*,	due to a disengagement of this gas, which, if the solution is concentrated, takes place with effervescence, but there is no deposite of sulphur formed (See Table I., E, Observation *h*, Table VII., Observation *f*)
SO_3	*Odour of SO_2*,	due to a disengagement of this gas, precisely as with HCl, but with SO_3 the odour of the gas evolved (SO_2) is much more distinct. No deposition of sulphur takes place. (See Table I., C, Observations *l*, *m*; Table VII., Observation, *f*)
NO_5 . .	*Odour of SO_2*,	if the solution is concentrated, and no heat is applied, but by boiling,
	Ruddy fumes	of NO_4 are evolved, and there is formation of SO_3, owing to which BaO,NO_5 produces then a white precipitate of BaO,SO_3 insoluble in water and in acids. (See Table I., C, Observations *n*, *o*.)
KO,MnO_3		A solution of green manganate of potash, reddened by a few drops of SO_3, becomes instantly
	Decolorised	by a solution of a sulphite, or of

		sulphurous acid. This is a very delicate test, but is not a characteristic one, since the acids in *ous*—phosphorous acid for example—and muriatic acid decolorise also the solution of manganate and of permanganate of potash. (See Table VII., Observation *g*.)
BaCl or CaCl	*White*	precipitate (BaO, SO_2, or CaO, SO_2), soluble in HCl. If, however, the solution has been left exposed to the air even for a short time, a white precipitate (BaO, SO_3) insoluble in water and in acids will be produced proportionate to the quantity of SO_2, which may have thus been transformed into SO_3 by combining with the oxygen of the air.
PbO, NO_5	*White*	precipitate (PbO, SO_2), soluble in cold NO_5, by boiling, the SO_2 becomes oxydised, and an insoluble precipitate (PbO, SO_3) is the result
AgO, NO_5	*White*	precipitate (AgO, SO_2) in neutral solutions soluble in a large excess of sulphite, while boiling it turns black, owing to a reduction of silver. The liquor afterwards contains SO_3. (See Table VII., Observation *h*.)
SnCl+HCl	*Brown*	deposite surrounding a crystal of pure SnCl after it has been dropped into a solution to which an excess of HCl has been added.
KO, NO_5+ BaO, NO_5		Thrown into fused nitre, sulphites become converted into sulphates, and if the mass be now dissolved and tested with BaO, NO_5, a
	White	precipitate (BaO, SO_3), insoluble in water and in acids will be produced.
I+alcohol		If an alcoholic solution of iodine is poured into that of a sulphite, the iodine solution is

Decolorised.	The solution should not contain free HCl. The decolorisation is due to a formation of a compound of sulphurous acid and iodine (iodosulphuric acid), which is colourless.

The most characteristic tests are the odour of SO_2 evolved when treated by SO_3 without sulphur being deposited. (See Table VII., Observation *h.*)

The brown colouring produced by a crystal of SnCl.

The white precipitate produced by BaCl, *soluble* in water and in acids.

The white precipitate insoluble in water and in acids, produced in the solution by BaO,NO_5 after fusion with KO,NO_5.

The white precipitate produced by AgO,NO_5, turning black or brown by boiling.

The solution of the green manganate of potash is very delicate and characteristic when the operator has ascertained the absence of the substances which may produce the same reaction.

The solution of iodic acid and starch is also a good test. (See Table XXVII., A, Observation *d.*)

TANNIC ACID, OR TANNATES.

($C_{18}H_5O_9$, 3HO, or $\overline{Qt}$,3HO)

Tannic acid, also called quercitannic acid, gallotannic acid, and tannine, is an odourless, white, solid, but uncrystallisable substance, of a very astringent taste, very soluble in water, less so in alcohol, and still less so in ether. Its aqueous solution slightly reddens litmus paper, but it rapidly absorbs the oxygen of the air, and becomes converted into gallic acid (C_7HO_3,3HO), carbonic acid being disengaged pending the

transformation. Pure and dry tannic acid or tannine is not altered by exposure, and behaves sometimes like an acid, and sometimes like a base; it combines with various mineral acids, forming with them white compounds which are insoluble in acids, but very soluble in water. Tannic acid does not produce any precipitate with organic acids. When boiled with dilute SO_3, or with HCl it is converted into gallic acid and glucose. The composition of tannic acid is as yet very doubtful.

TESTS AND REACTIONS

NH_3 . . .	*Brown* . .	colour.
CaCl. . . .	*White* . .	precipitate, almost insoluble in water.
NH_4Cl . .	*Turbidness.*	or precipitate
Fe_2Cl_3 . .	*Black* . .	colour
FeO, SO_3 . .	*Nothing.* .	at first, but by exposure the liquor assumes an intensely
	Bluish-black	colour (ink).
Solution of gelatine	*Whitish flocculent.* .	precipitate, which, if an excess of tannic acid is present, agglomerates by boiling into a viscid elastic mass

TANTALUM, OR COLUMBIUM.

(Ta.)

Tantalum is found in the state of tantalic acid, combined with various other oxydes, in a few rare minerals only (tantalite—yttrotantalite), and is, therefore, a very rare metal, which is black when in powder, but which by friction acquires a grey metallic lustre. It is infusible, even in the blast-furnace; but, heated in contact with the air, it becomes converted into tantalic acid. It is insoluble in nitric acid, and, even in aqua regia; solution of potash has no action upon it, but, fused with the alkalies or their carbonates, it is converted into a tantalate of alkali.

TANTALIC ACID.

($Ta_2 O_3$.)

Tantalic acid is white, insoluble, infusible, and indecomposable by heat. It dissolves in HFl, but is insoluble, or almost so, in all other acids. It is quite soluble in a solution of KO, from which it is reprecipitated by HCl, and the other acids, even by CO_2. Fused with microcosmic salt, before the blowpipe, it yields a colourless, transparent bead; and, with borax, the glass produced is limpid, colourless also, but it becomes milky as soon as it cools. These two blowpipe reagents distinguish it readily from SiO_3.

TARTARIC ACID AND TARTRATES.

($C_8H_4O_{10}$, 2HO, or $\bar{T}$, 2HO)

Tartaric acid is a bibasic acid, that is to say, it is an acid which requires two equivalents of base to produce neutral salts. Its crystals, often of large size, are colourless, and transparent, and have the form of oblique rhombic prisms, which are inalterable in the air. Tartaric acid, however, is generally met with in commerce in the form of a white crystalline powder. This acid has an agreeable, sour taste, it is very soluble in water, and in alcohol. It is easily decomposed by oxydising agents into carbonic and formic acids. The aqueous solution becomes mouldy by keeping.

Tartrates may be the result of a combination of tartaric acid with 2 equivalents of a same base; as, for example, (2CaO),$\bar{T}$; or by 2 equivalents of different bases; as, for example, KO, NaO, $\bar{T}$; or by 1 equivalent of base + 1 equivalent of water, the latter replacing 1 equivalent of base; as, for example, KO, HO, $\bar{T}$ (cream of tartar): this then constitutes what are called acid tartrates. The compounds called *emetics*,

are those in which tartaric acid is saturated by one base containing 1 equivalent of oxygen + one base containing 3 equivalents of oxygen, as, for example, $KO,Sb_2O_3,\bar{T}$,—or in which the second equivalent of base is replaced by 1 equivalent of a weak acid, such as, for example, $KO,BO_3,\bar{T}$.

The tartrates of alkalies are soluble in water, and all the tartrates are soluble in HCl, and in NO_5; they interfere with, or altogether prevent, the precipitation of peroxyde of iron, alumina, and protoxyde of manganese, by alkalies.

The tartrates, and tartaric acid, are decomposed by exposure to a red heat, an odour of burnt bread, or burnt sugar, being evolved at the same time. (See Table XXII., A, Observations *e*, *h*.)

TESTS AND REACTIONS.

NH_3	*white*. . .	In concentrated solutions of $\bar{T}$, 2HO, bulky precipitate, soluble in a slight excess of the reagent and in water. Ammoniacal salts prevent this precipitate (See Table VIII, Observation *a*.)
CaCl . . .	*White* . .	precipitate (in neutral solutions), almost insoluble in water, but soluble in sal-ammoniac, the presence of which retards the production of the precipitate, according to the state of concentration of the liquor. The liquor should be violently shaken after the addition of the reagent. This white precipitate is soluble in an excess of a strong solution of caustic KO, but is immediately reproduced by boiling. This reaction is *very characteristic*.
CaO+aq.	*White* . .	*Lime water* produces a precipitate in solutions of tartaric acid or of tartrates, but an excess of the reagent must be employed, since this white precipitate, which is tartrate of lime, is soluble in free $\bar{T}$,

2HO and other acids, and likewise, as we just said, in sal-ammoniac.

$KO,\bar{A}+\bar{A},HO$ . . . . Acetate of potash, or other solutions of salts of potash, produce a

White crystalline . precipitate ($KO,HO,\bar{T}$, bitartrate or acid tartrate of potash), sparingly soluble in water. The mixture should be violently shaken, and alcohol increases the delicacy of the test. If *sulphate of potash* is used instead of *acetate of potash*, an excess must be avoided, since the precipitate would be redissolved by it.

$Fe_2Cl_3.$ $+KO$ . . . . . KO or NH_3 produce, as will be recollected, a red precipitate of peroxyde of iron in solutions of this oxyde; but if tartaric acid, or a tartrate is present in sufficient quantities,

No precipitate will be produced by these reagents A precipitate takes place only when the proportion of tartaric acid or of tartrate is too small.

CaO,SO_3 . *No precipitate* in solutions of $\bar{T}$, 2HO, or of tartrates, and this is a character by which tartaric acid may be distinguished from racemic acid and racemates.

AgO,NO_5 . . *White* . . curdy precipitate, insoluble in HO, soluble in NH_3

The best tests are CaCl and $KO,\bar{A}+\bar{A},HO$.

TELLURIUM.

(Te.)

Tellurium is a rare metal or semi-metal, of a white colour, resembling silver, but not so lustrous. It is brittle, easily fusible, and it volatilises without alteration, when heated out of the contact of the air, but with the contact of the air, it becomes converted into tellurous acid, which

sublimes in the form of a white smoke. Its specific gravity is 6·26. Heated before the blowpipe upon charcoal, it imparts a blue colour to the flame. It is very soluble in *nitric acid* and in *aqua regia*, the solution thus obtained contains tellurous acid (TeO_2). Tellurium is insoluble in HCl, and in solution of *potash*. Two compounds of this metal with oxygen are known, namely, tellurous (TeO_2) and telluric (TeO_3) acids.

TELLUROUS ACID.

(TeO_2.)

Tellurous acid in the state of a white powder, or of octahedric crystals, is only slightly soluble in water and in acids, but dissolves readily in the alkalies.

TESTS AND REACTIONS.

HS	*Brown*	(in acid solutions), immediate precipitate, resembling much that produced by that reagent with protosalts of tin.
NH_4S	*Brown*	precipitate (in neutral solutions), readily soluble in an excess of the reagent.
NaO,SO_3	*Black*	(In acid solutions) precipitate of metallic tellurium especially by heating the liquor
Infusion of galls	*Dingy*	precipitate
FeO,SO_3	*Black*	precipitate of metallic tellurium.
SnCl	*Black*	ditto.

TELLURIC ACID.

(TeO_3.)

Telluric acid is soluble in water, though slowly; it has a metallic non-acid taste; it is soluble also in alcohol. The

tellurates exposed to a red heat fuse, and are converted into tellurites, oxygen being disengaged. HCl dissolves tellurates without decomposition, except heat be applied, in which case chlorine is disengaged, and on adding water a white precipitate of TeO_2 is produced.

TESTS AND REACTIONS.

HS	*Nothing*. .	except after a very long time.
NaO,SO_2 . .	*Black* . .	precipitate of metallic tellurium by boiling.
Infusion of galls	*Nothing*. .	

THORINIUM.

(Th.)

Thorinium is a rare metal which resembles Aluminium. It is slowly attacked by acids, but is not oxydised by water either hot or cold; the alkalies have no action upon it. HFl dissolves it rapidly. There is only one compound of oxygen and thorinium, namely, thorina (ThO).

THORINA.

(ThO.)

Thorina is a white oxyde, insoluble in water, and is the heaviest of the earthy oxydes, its specific gravity being 9·402. Its hydrate ThO,HO is insoluble in the alkalies, but is soluble in their carbonates, and in all the acids, except after strong ignition. The salts of thorina are colourless, and have an astringent taste.

TESTS AND REACTIONS.

KO or . . NH_3 . .	*White* . .	gelatinous precipitate, insoluble in excess.
KO,CO_2 or . NaO,CO_2 or NH_4O,CO_2 .	*White* . .	precipitate, soluble in excess. (See Table XXXI., Observation *e*.)

(2Na,O,)HO,PhO_5 or $\bar{O}$, HO, or K_2Cfy	*White*	precipitate.
KO,SO_3	*White*	precipitate (ThO, KO,SO_3).
NH_4S	*White*	precipitate (ThO,HO).

Thorina differs therefore from Al_2O_3 because it is insoluble in KO;—from yttria because it forms a double sulphate with KO,SO_3, insoluble in a saturated solution of the latter reagent;—from magnesia, because the solution of its salts previously acidified, are precipitated by NH_3.

TIN.

(Sn.)

Tin is a metal which is nearly as white as silver, but with a slight tinge of yellow, especially after exposure, when rubbed between the fingers it has a disagreeable smell It is very malleable and soft, and when bent or twisted, it emits a crackling sound. It is one of the least elastic metals, wherefore it is not sonorous. Its specific gravity is 7.285, and it is not increased by hammering. Tin melts at 442° Fahr., and does not volatilise at any temperature, however high. Tin is scarcely altered by exposure; but when heated in contact with the air, it soon becomes converted, first into SnO, and then into SnO_2.

Dilute SO_3 has scarcely any action upon tin in the cold; but by boiling, it is dissolved, hydrogen is disengaged, and SnO,SO_3 formed. Boiling sulphuric acid dissolves tin also; SO_2 is disengaged.

Concentrated HCl dissolves tin rapidly, and converts it into SnCl, and hydrogen gas of a fetid odour is disengaged. The action of dilute HCl on tin is very slow.

Concentrated NO_5 attacks tin with the help of heat, and converts it into SnO_2 (or hydrated metastannic acid, Sn_5

O_{10}, 10HO) which remains perfectly insoluble in the acid. Very dilute NO_5 attacks tin also, but very slowly.

Aqua regia dissolves tin rapidly, the result being $SnCl_2$; but if the aqua regia contains an excess of NO_5, then SnO_2 is produced.

Solutions of the *pure alkalies* dissolve tin, hydrogen being disengaged, and a soluble metastannate of the alkali produced.

PROTOXYDE AND PROTO-SALTS OF TIN.

(SnO)

SnO is a greyish black or brown powder in the dry state. Its hydrate (SnO,HO) is white, and, like all hydrates, is more soluble in acids than the anhydrous oxyde. NO_5 converts it into SnO_2, insoluble in that acid. SnO when set fire to burns in contact with the air like tinder, and is thereby converted into SnO_2.

The protosalts of tin are colourless; they spontaneously absorb oxygen, and become converted into salts of peroxyde; for this reason the protosalts of tin are powerful deoxydising agents. SnCl is remarkable in this respect, and on that account forms a valuable reagent, for it immediately reduces a great number of oxydes, such as oxydes of antimony, of zinc, of mercury, of silver, arsenious and arsenic acids. It also converts the peroxyde of copper, of iron, of manganese, into suboxyde of copper, protoxyde of iron, and protoxyde of manganese, &c. The aqueous solution of SnCl is always milky, which is owing to the production of a basic chloride of tin, which is insoluble (See Table IX., Observation *i.*), whilst a neutral and an acid chloride remain in solution. All the salts of tin have an acid reaction.

TESTS AND REACTIONS.

HS . . . . *Dark brown* precipitate (SnS) in neutral and in acid solutions, soluble in alkaline

		sulphurets, and in solution of KO. The alkaline sulphurets should, however, contain an excess of sulphur, or else a very large excess of reagent is required (See Table XVI., Observation *d*, Table XV., Observation *h*; Table XXIII., Observation *q*)
NH_4S . . .	*Dark brown*	precipitate (SnS), difficultly soluble in an excess of the reagent. If, however, the NH_4S contains an excess of S, the precipitated sulphuret is readily dissolved by it. Acids reprecipitate the dissolved sulphuret in the state of SnS_2 mixed with sulphur
KO . . .	*White* .	precipitate (SnO, HO), soluble in a great excess of the reagent.
NH_3 . .	*White* . .	precipitate (SnO, HO), insoluble in an excess of the reagent.
KO, CO_2 . .	*White* .	precipitate (SnO, HO), insoluble in an excess of KO, CO_2.
HgCl .	*White* . .	precipitate (Hg_2Cl), becoming *greyish* after some time.
$AuCl_3$ . .	*Purple* . .	precipitate (purple of Cassius). The liquor, however, must contain a little perchloride of tin at the same time (See Table II, B, Observation *g*)
$PtCl_2$		in *dilute* solutions a
	Reddish-brown	flocculent precipitate resembling Fe_2O_3. In *more concentrated solutions* the liquid assumes an almost
	Blood-red .	colour, *but no precipitate takes place*, nor is a precipitate produced by subsequently diluting the liquid which has a colour resembling that produced by sulphocyanogen in persalts of iron. (See Table IV., Observation *z*.)

A bar of zinc precipitates tin from its solution in the metallic state.

Blowpipe.—Mixed with NaO,CO_2, or with one part of NaO,CO_2 and two parts of KCy, and heated in the inner flame of the blowpipe on a charcoal support, bright, metallic, malleable grains of tin easily oxydisable in the outer flame, are obtained. (See Table I., B, Observation *j.*) These grains should be dissolved in HCl, and tested by $AuCl_3$. (See Table II., B, Observation, *g.*)

The most characteristic tests for the protosalts of tin, are HS,—NH_4S,—$AuCl_3$, and $PtCl_2$. (See also Table XVI., Observation *e.*)

PEROXYDE AND PERSALTS OF TIN.

(SnO_2.)

SnO_2 produced by the action of NO_5 is a white powder, insoluble in water, in nitric, and in dilute SO_3, but concentrated SO_3,HO dissolves it readily.

SnO_2 produced by precipitation, is *white* also, but it is soluble in the alkalies and in acids; if however it be ignited, it then becomes quite as insoluble as when produced by the action of NO_5 as just said. After ignition it often has a *yellowish colour.*

A dilute solution of $SnCl_2$ is not decomposed or precipitated by NO_5, but when the liquid has attained a certain degree of concentration, then nitrous acid fumes are evolved, and a precipitate of *insoluble* SnO_2 takes place.

The persalts of tin are colourless.

TESTS AND REACTIONS.

HS . . . . *Yellow* . . precipitate (SnS_2) which, however, requires some time to be produced, and augments by standing. The precipitation is hastened by boiling the liquid after complete saturation with HS. It is soluble in NH_4S,

		and other alkaline sulphurets, and likewise in KO, and in HCl. (See Table XVI., Observation *d.*)
NH_4S . . .	*Yellow* . .	precipitate (SnS_2), immediately soluble in an excess of the reagent.
KO	*White* . .	bulky gelatinous precipitate (SnO_2, HO), soluble in an excess of the reagent.
NH_3 . . .	*White* . .	precipitate (SnO_2,HO), bulky; somewhat soluble in an excess of KO, less soluble in an excess of NH_3, and still less in an excess of an alkaline carbonate, the solution, however, remains milky in all cases, and deposits again after some time a *white bulky precipitate.* The precipitate is readily soluble in all acids.

Blowpipe.—Mixed with NaO,CO_2, or better still, with one part of NaO,CO_2 and two parts of KCy, and heated in the inner flame of the blowpipe upon a charcoal support, a white metallic and malleable bead is obtained, easily oxydised when heated in the outer flame. (See also Table XVI., Observation *c.*)

The most characteristic test for the persalts of tin is HS, and the blowpipe.

TITANIUM.

(Ti.)

Titanium, when in compact masses, has the appearance of copper; it is very bright, and hard enough to scratch quartz. When in powder it looks black, but friction imparts to it the metallic lustre. It is a brittle metal, the density of which is 5·9. It is insoluble in *sulphuric, nitric* and *hydrochloric acids,* and even in *aqua regia,* but fusion with a mixture of nitrate and of carbonate of potash converts it into titaniate of potash.

TITANIC ACID.

(TiO_2)

Titanic acid is a white powder, which turns yellow when heated, but becomes white again on cooling. It is insoluble in all acids, except in boiling concentrated SO_3HO. Ignited with the alkalies or their carbonates, it combines with them and forms titaniates of alkalies. The excess of alkali is dissolved by water, but the titaniate is left behind, this titaniate being insoluble in water, but soluble in HCl with the help of heat. Titanic acid is distinguished from other substances because it is precipitated from its acid solutions by *boiling;* the boiling, however, must be continued for a long time; the titanic acid then falls down in the form of a white powder. A bar of zinc plunged in a muriatic acid solution of a titaniate of alkali, *imparts a blue colour to the liquid;* or if thrust into the precipitate produced by boiling, *the precipitate becomes blue,* the colour beginning from the point of contact with the zinc. Fused with microcosmic salt before the blowpipe, a *blue bead* is produced in the *interior* flame only. (See Table XXIX., Observation *g*.)

TUNGSTEN. (WOLFRAM.)

(W.)

Tungsten is a metal of a dark grey colour, non-volatile, almost infusible, and very hard. Its specific gravity is 17·6.

Heated in contact with the air it becomes converted into tungstic acid (WO_3).

SO_3 and HCl have no action upon tungsten

NO_5 or *aqua regia* convert it into tungstic acid.

Nitrate of potash and the alkalies convert it into tungstate of alkali.

TUNGSTIC ACID.

(WO_3.)

It is a yellow powder, soluble in the alkalies, but insoluble in water. It reddens litmus paper only slightly. The tungstates of alkalies are soluble in water; but these solutions are precipitated *white* by HCl, by NO_5, and by SO_3; heat renders this precipitate *yellow*, and it is not soluble in an excess of the acid by which it was produced, as is the case with the molybdates, which are also precipitated under similar circumstances. A bar of metallic zinc plunged into the precipitate produced by adding an acid (but not sulphuric acid) to the solution of a tungstate, and without separating the precipitate from the liquor, imparts to the precipitate a fine blue colour. Fused with microcosmic salt a fine blue bead is obtained in the interior flame.

URANIUM.

(U.)

Uranium is a black powder or a white malleable metal, which is not oxydised by exposure to the air, or by the contact of water. It burns vividly at a moderate heat, and is soluble in the dilute acids, hydrogen being disengaged, and a green salt of uranium produced. The salts of protoxyde of uranium are green; those of peroxyde of uranium are yellow.

TESTS FOR THE PROTOSALTS OF URANIUM.

KO or NH_3 . .	*Brownish-black* precipitate (UO,HO), which becomes *yellow* by exposure, because it becomes converted into peroxyde.
$\bar{O}$,3HO . . .	*Greenish-white* precipitate (UO,$\bar{O}$,3HO).
NH_4S . . .	*Black precipitate* (US).

The green salts of uranium have a great affinity for oxygen, and they reduce the salts of gold and of silver. NO_5 converts the protosalts of uranium into persalts.

TESTS FOR THE PERSALTS OF URANIUM.

Alkalies . .	*Yellow* . .	precipitate (KO, U_2O_3).
NaO, CO_2 . .	*Yellow* . .	precipitate, soluble in an excess of NH_4O, CO_2, reprecipitated by boiling.
K_2Cfy . . .	*Reddish-brown* precipitate.	
NH_4S . . .	*Brownish-yellow*, almost insoluble in excess	

YTTRIUM.

(Y.)

Yttrium is a very rare metal, the earth yttria, from which it is obtained, existing only in a very few rare minerals, such as *gadolinite*, which is a combination of yttria, silica, glucina, the oxydes of cerium and iron.

YTTRIA.

(YO.)

Yttria is the only combination of yttrium with oxygen; it is an earth which is more slowly dissolved by concentrated than by dilute acids.

TESTS FOR THE SALTS OF YTTRIA.

KO or NH . . .	*White* . .	precipitate (YO,HO), insoluble in excess.
KO, CO_2 or NaO, CO_2 . .	*White* . .	precipitate (YO, CO_2), *entirely soluble in an excess* of the reagent. (See Table XXXI., Observation *e.*)

$\bar{O}$, HO . . .	*White* . .	precipitate (YO, $\bar{O}$), in slightly acid solutions.
KO, SO . .	*White* . .	precipitate (KO, YO, SO_3), after some time; the precipitate is sparingly soluble in a large quantity of water.

ZINC.

(Zn.)

Zinc is a metal of a bluish-white colour, which slowly tarnishes by exposure. Its specific gravity is 6·8; its texture is lamellar; it is rather brittle in the cold, but can be easily laminated at about 250_0 Fahr., and afterwards it retains its malleability. At about 400 it can be pulverised, and it melts at about 800° Fahr. At a white heat in close vessels it evaporates and distils over. If the air be then admitted, it burns with a beautiful greenish white flame, and a production of thick solid white fumes of oxyde of zinc. Solution of *potash*, of *soda*, or of *ammonia*, dissolves zinc, producing zincates of alkalies, and hydrogen is disengaged. Metallic zinc is soluble in the dilute acids.

OXYDE AND SALTS OF ZINC.

(ZnO.)

Pure oxyde of zinc is of a beautiful *white* colour. When heated, and whilst hot, it becomes yellow, but it re-assumes its original white colour after cooling. Sometimes, however, when it has been strongly heated, it retains after cooling a slight tinge of yellow, possibly from the presence of a trace of iron.

Hydrated oxyde of zinc (ZnO,HO) is white, and when recently precipitated is rapidly dissolved by acids, even

when very dilute, but drying it simply in the air renders it of more difficult solution in acids, and it is not then dissolved by the alkalies unless heat is applied.

The salts of zinc are colourless, and have a styptic taste, and all those which are soluble in water have an acid reaction on litmus paper; they are decomposed by a red heat.

TESTS AND REACTIONS.

NH_4S		in neutral and in alkaline solutions
	White . .	precipitate (ZnS), insoluble in any excess of the reagent, in KO, and in NH_3, but it is readily dissolved by acids. A *white* precipitate produced by a soluble sulphuret in a colourless and *strongly alkaline* solution can hardly be anything else than ZnS. (See Table V., Observation *f*; Table XVIII., Observations *d, i*)
HS	*No precipitate*	in acid solutions, but if the *acid is a weak acid*, acetic acid, for example, or if an excess of $NaO, \bar{A},HO$ be added, the whole of the zinc may be precipitated in the form of a white precipitate (ZnS).
KO or NH_3 . .	*White* . .	precipitate (ZnO,HO) of a gelatinous appearance, soluble in an excess of the reagent.
NH_4O,CO_2 .	*White* . .	precipitate ($ZnO,(3HO) + 2ZnO,CO_2$), soluble in an excess of the reagent.
KO,CO_2 . .	*White* . .	precipitate (ZnO,CO_2), insoluble in an excess of the reagent, but if ammonical salts are present, *no precipitate* is produced.

Blowpipe.—Heated on charcoal, with NaO,CO_2, in the reducing flame, volatilisation takes place, and the charcoal

becomes covered with a yellow incrustation, which on cooling becomes white. By moistening the above-mentioned precipitate with CoO,NO_5, and then heating the mass before the blowpipe, it assumes a beautiful *emerald green colour.*

The most characteristic tests for zinc are NH_4S, which in strongly alkaline and clear solutions produces a white precipitate (ZnS), difficult to filter, and the emerald-green colour which ZnO assumes when moistened with CoO,NO_5 and heated before the blowpipe.

ZIRCONIUM.

(Zr.)

Zirconium is a rare metal, of a black colour, to which friction imparts only an imperfect metallic lustre. It is difficult to fuse. Heated in contact with the air, at a temperature much below redness it burns, and becomes converted into zirconia (Zr_2O_3). Fusion with nitre, or with KO,ClO_5, oxydises it only slowly, but by fusing it with the free or carbonated alkalies, or even with borax, it is rapidly converted into oxyde of zirconium (Zr_2O_3).

$$\left.\begin{array}{l} HCl \text{ or . .} \\ SO_3 \text{ or . .} \\ NO_5 \text{ or. .} \\ \text{Aquaregia} \end{array}\right\} \text{. . . . have scarcely any action upon it.}$$

The only acid by which it is rapidly dissolved, is HFl.

ZIRCONIA.

(Zr_2O_3.)

Zirconia is a white, tasteless, and inodorous earth, insoluble in water, and which, after ignition, is hard enough to scratch glass; it is then insoluble in all acids, except concentrated SO_3,HO. It is slightly soluble in the alkaline carbonates.

TESTS FOR ZIRCONIA.

KO or NH_3 . . .	*White* . .	precipitate (Zr_2O_3, HO), insoluble in an excess of the reagent; soluble to a certain extent in alkaline bi carbonates.
Alkaline carbonates and bi-carbonates . . .	*White* . .	bulky precipitate (Zr_2O_3, CO_2), sparingly soluble in an excess of the reagent. (See Table XXXI., Observation *e.*)
KO, SO_3 . .	*White* . .	precipitate, after some time; soluble in a large excess of HCl.
NH_4S . . .	*White* . .	precipitate (Zr_2O_3, HO). If $\overline{T}$,2HO is present no precipitate is produced.

A

DICTIONARY OF TESTS & REAGENTS;

INDICATING

THEIR PREPARATIONS FOR THE LABORATORY,
THE MEANS OF TESTING THEIR PURITY,
AND THEIR BEHAVIOUR WITH SIMPLE SUBSTANCES
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A

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ACETATE OF AMMONIA.

($NH_4O,\bar{A}$.)

ACETATE OF POTASH OR OF SODA.

($KO,\bar{A}$ or $NaO,\bar{A}$)

These acetates are prepared respectively for the laboratory by saturating acetic acid with ammonia; or by dissolving carbonate of potash or of soda in hot water, and whilst the solution is kept hot, adding gradually pure acetic acid to exact neutralisation. The liquor may then be evaporated to dryness, or, if it is acetate of soda, to the crystallising point. As neither acetate will keep in solution, it is better to dissolve a portion of the dry acetates when wanted; the acetate of ammonia can always be readily prepared at once.

Acetate of potash is chiefly used to detect *tartaric acid*, a precipitate of bitartrate of potash being produced, which is insoluble in the free acetic acid resulting from the reaction.

As a test for tartaric acid, acetate of potash is better than any other salt of that base; the addition of alcohol, in which bitartrate of potash is much less soluble than in water, augments the delicacy of this test.

Either acetate is used for the detection of *phosphoric acid* (see Phosphoric Acid, and Table IX., Observation *n*); for the detection of *arsenic* (see Arsenic, in the other Dictionary); and for that of the earthy oxalates (see Table V., Observation *r*).

Acetate of soda is used also, instead of acetic acid, for the purpose of substituting acetic acid to the mineral acids which may exist in the free state in a liquor, and which combining with the soda of the acetate form a nitrate, a sulphate, or a chloride of sodium, whilst free acetic acid is liberated.

From solutions thus treated with acetate of soda, or any other soluble acetate, zinc may be precipitated by HS.

ACETATE OF BARYTES.

($BaO,\bar{A}$.)

Acetate of Barytes is a very soluble salt, which behaves with substances like the other soluble salts of Baryta.

When treated by water, the solution which it yields is sometimes not perfectly clear, and it is likewise occasionally contaminated by HCl, which renders it unfit for use as a special test for separating magnesia from the alkalies; the presence of HCl is easily detected by treating the aqueous solution of the salt with a little nitric acid, and then adding AgO,NO_5, which will then produce a white precipitate.

Acetate of Barytes, like Baryta water, and like the nitrate and chloride of the same base, is used for precipitating sulphuric acid, instead of the nitrate and chloride, in cases where it is necessary to avoid the presence of mineral acids, and also as a special reagent for separating magnesia from the alkalies. It is used also in separating oxyde of zinc from peroxyde of iron; the two oxydes being first dissolved in sulphuric acid, and acetate of Baryta added, when a current of HS passed through the solution will precipitate the zinc in the state of ZnS.

ACETATE OF LEAD (BASIC).

$(3PbO,\bar{A}.)$

Digest at a moderate heat 7 parts of finely pulverised litharge, 6 parts of ordinary or neutral acetate of lead (PbO,A) in 30 parts of water; keep the whole in a well-corked flask until the superincumbent liquid is quite clear, decant, and keep the clear liquor in a well-closed bottle, in which a few strips of perfectly pure lead are kept.

This reagent behaves with compounds like the neutral acetate, but it is especially employed for the detection of sulphuretted hydrogen, for which it is a more sensitive test than the neutral acetate. It is used also for obtaining certain alkaloids.—Its solution should restore the blue colour of reddened litmus paper, and produce a thick precipitate in a solution of gum arabic.

ACETATE OF LEAD (NEUTRAL).

$(PbO, \bar{A}.)$

Take acetate of lead of commerce, of good quality, and dissolve one part in weight of it in about 10 parts of water. PbO,NO_5 generally answers the same purpose Its behaviour with substances is as follows:—

REACTIONS.

Hydrochloric acid (chlorine—chlorides) .	*White* . .	precipitate; sparingly soluble in water.
Hydrobromic acid (bromides) .	*White* . .	precipitate; insoluble in water.

Hydrofluoric acid	*White* . .	precipitate.
Hydriodic acid (iodides, soluble)	*Orange yellow*	precipitate; soluble in hot water and in NO_5· the solution deposits fine spangles, of a gold colour.
Hydrocyanic acid	*White* . .	precipitate.
Chromic acid	*Yellow* .	precipitate; soluble in KO
Boracic acid	*White* . .	precipitate; soluble in acids.
Sulphuric acid	*White* . .	precipitate; almost insoluble in an excess of the reagent.
Phosphoric acid	*White* . .	precipitate.
Osmic acid	*Nothing;* . *Deep brown*	In neutral solutions but in basic solutions, or by adding ammonia, a precipitate is produced
Arsenic acid	*White* . .	precipitate.
Carbonic acid	*White* . .	precipitate.
Sulphuretted hydrogen	*Black* . .	precipitate.
Hydrosulphuret of ammonia (alkaline sulphurets)	*Black* . .	precipitate
Bromic acid	*White* . .	precipitate; soluble in a great excess of water.
Iodic acid	*White* . .	precipitate; soluble in NO_5.
Sulphurous acid (sulphites)	*White* . .	precipitate; soluble in cold NO_5: by boiling, nitrous acid fumes are evolved, and sulphate of lead precipitated.

Molybdates of alkali.	*White* . .	precipitate.
Tungstates .	*White* . .	precipitate.
Oxalic acid .	*White* . .	precipitate, abundant; sparingly soluble in free $\overline{O},HO$; soluble in NO_5; insoluble in NH_3.
Tartaric acid .	*White* . .	precipitate; insoluble in HO; soluble in NH_3.
Citric acid .	*White* . .	precipitate, abundant; sparingly soluble in NH_3.
Malic acid .	*White* . .	precipitate, abundant; becoming viscid by boiling.
Succinic acid .	*White* . .	precipitate; insoluble in an excess of the reagent.
Benzoic acid .	*White* . .	precipitate; not immediately, except in benzoate of potash and of soda.

ACETATE OF POTASH. (See ACETATE OF AMMONIA.)

ACETATE OF SODA. (See ACETATE OF AMMONIA.)

ACETIC ACID.

($C_4H_3O_3,HO$ or $\overline{A},HO$.)

The acetic acid used as a reagent must be pretty strong. It must not contain any HCl—SO_3—SO_2—NO_5, nor metallic oxydes, especially lead and copper; it should therefore evaporate without leaving any residue, and not be precipitated by AgO,NO_5, nor by BaCl or BaO,NO_5, even after boiling with NO_5 (which would convert any sulphurous acid which might be present into sulphuric acid) nor by HS, nor must it decolorise solution of indigo.

It may be easily prepared as follows:—Mix together 12 parts in weight of concentrated sulphuric acid with an equal

weight of water, stir the mixture well, and let it cool. On the other hand, triturate together 100 parts in weight of neutral acetate of lead, and 30 parts, also in weight, of effloresced sulphate of soda, put the pulverised mass into a retort, and then pour upon it the cold mixture of sulphuric acid and water above alluded to. Connect the retort with a refrigerator or Liebig's condenser, and distil the whole to dryness by means of a sand-bath. Or else reduce to powder 3 parts of dry acetate of soda, pour upon it 9·7 parts of pure and very concentrated SO_3,HO, and at once connect the whole with a Liebig's condenser; one-eighth part of the acetic acid will spontaneously distil over by the heat developed by the mixture itself, and the rest of the acetic acid will pass over upon applying a gentle heat to the mixture. The distillate will probably require purification, which is accomplished by submitting it to a second distillation.

If it has been found that it is contaminated by SO_2, the best way will be to digest in it some pure oxyde of lead (PbO_2), which will transform it into SO_3, and combine with it; the superincumbent liquor may then be carefully decanted, and further purified by redistillation.

Acetic acid is sometimes used for acidifying liquors when the use of mineral acids would be objectionable, or to precipitate by HS certain substances, oxyde of zinc for example, the solutions of which in the mineral acids cannot be precipitated by that reagent. The acetic acid fortis of commerce is good enough for such a purpose.

Acetic acid is used also to separate phosphate from oxalate of lime; the former being soluble, and the latter insoluble in that acid.

Acetic acid produces a white precipitate in solutions of tungstic acid, and the precipitate is insoluble in an excess of the reagent; time does not render it *yellow*, as is the case with the white precipitate produced by HCl or SO_3 in solutions of the tungstates.

ALCOHOL.

($C_4H_6O_2$.)

Both absolute alcohol and spirits of wine are used as tests; but rectified spirits of wine of specific gravity 0·835 or 0·84, provided it is pure, is generally good enough for most analytical purposes.

Absolute alcohol is obtained by mixing pure rectified alcohol with about half its weight of newly burnt quick-lime in coarse powder, leaving the mixture at rest for a few days, and distilling the alcohol therefrom by means of the water-bath. The distillate thus obtained, however, is not quite free from water, and it may be necessary to mix it a second time with quick-lime, and to redistil it. The alcohol is known to have lost its water by taking its specific gravity, which should then be from 0·792 to 0·800, and also by putting a certain quantity of anhydrous sulphate of copper in contact with it, in a well stoppered flask; if the alcohol is free from water, the anhydrous sulphate of copper will remain white; but in the contrary case, it will become blue.

Both absolute alcohol and rectified spirits of wine should completely evaporate without the slightest residue, nor should they have any action upon litmus paper.

Alcohol, or spirits of wine, is used principally as a test for boracic acid and strontia; for deoxydising chromic acid and converting it into sesquioxyde of chromium; for promoting the formation of the yellow precipitate produced by $PtCl_2$ in solutions of potash and in that of sulphate of lime; and likewise for the separation of chloride of calcium from chloride of strontium. It is, however, more frequently used in the analysis of organic than of mineral substances.

REACTIONS.

Strontia . . *Crimson flame* when mixed with alcohol and inflamed.

Boracic acid .	*Green flame*	when mixed with alcohol, and inflamed. (See Table XXII.—A, Observation *b.*)
Osmic acid .	*Nothing* .	at first; but a
	Black . .	precipitate is produced after a time.
Sulphate of lime . .	*White* . .	precipitate (in solutions of,) especially by adding alcohol.
Potash		mixed with alcohol, burns with a
	Purple flame;	the smallest quantity of a salt of sodium interferes.
Soda . . .	*Yellow flame*	when mixed with alcohol and inflamed.

AMMONIA.

(NH_3 or NH_4O)

Ammonia is one of the reagents most frequently in use. Ammonia fortissima is a saturated, or almost saturated, aqueous solution of gaseous ammonia, it has a specific gravity of about 0·89 and contains about 29 per cent. of ammonia. It should be kept in well-stoppered bottles. For ordinary purposes this saturated solution may be diluted with three or four times its bulk of distilled water.

Ammonia may generally be found in commerce in a state of great purity. It may, however, be easily prepared by mixing, in a pretty large flask or matrass, about equal parts of slacked lime and of sal-ammoniac, both pulverised, with just enough water to moisten the mixture, so that on shaking the mass it agglomerates into small lumps The matrass is then connected with a washing bottle containing a very small quantity of water, through which the gas is passed, and the washing bottle is provided with a disengagement-tube plunging into a flask or bottle full of distilled water. The joints being made tight, if necessary with wax, the large flask containing the mixture is put upon a sand-bath, and heat being cautiously applied, the disengagement of gas takes places in a quick and regular manner. As ammonia is lighter than water, the dis-

engagement-tube should plunge to the bottom of the bottle full of distilled water, which should be placed in a jar or basin of cold water in order to keep it cool, for the absorption of the gas by the water is attended with a disengagement of heat The matrass containing the mixture of sal-ammoniac and lime should be provided with a safety-tube. The operation is continued as long as bubbles of gas are emitted.

Aqueous solution of ammonia should be as clear as water, the slightest tinge of brown indicates the presence of organic matter. It should leave no residue by evaporation, though an exceedingly small carbonaceous stain is generally left with most samples found in commerce and is of no consequence, it is only when the carbonaceous residue so left is weighable, that the reagent is unfit for analytical purposes.

Liquor ammoniæ is sometimes contaminated by NH_4O,CO_2, —NH_4Cl—CuO. Wherefore, after supersaturating a sample of it with NO_5, it should not be discoloured, rendered turbid, or precipitated by

Lime-water, which, in the contrary case, indicates		CO_2.
BaO,NO_5	ditto	SO_3.
AgO,NO_5	ditto	HCl.
HS	ditto	Metallic oxyde.

Aqueous solution of ammonia is principally employed for neutralising acid liquids, and for precipitating certain bases from their solutions, several of which, however, are redissolved by an excess of this reagent, which circumstance is turned to account for the separation of various bases.

The reactions of ammonia with the following substances are as follows:

REACTIONS.

Potash, Soda, Lithia,	*Nothing.*	
Baryta . . .	*Nothing.*	. for a long time, but after standing a considerable time (a few days f--

		example), a *white* precipitate (BaO, CO_2) is formed, which is due to an absorption of the CO_2 of the air. If the NH_3 contains NH_4CO_2, of course a precipitate is at once produced
Strontia .	*Nothing* .	for a long time; after a considerable time, a *white* precipitate (StO,CO_2) is produced, as with baryta.
Lime	*Nothing*	for a long time, but by exposure a *white* precipitate (CaO,CO_2) is produced.
Magnesia	*White* . .	precipitate (MgO,HO); almost entirely soluble in solution of NH_4Cl. If the solution be very acid, or contains a sufficient quantity of ammoniacal salts, nothing. (See Table V., Observation *a* Table VI., Observation *a*.)
Alumina .	*White* . .	bulky precipitate (Al_2O_3,HO), soluble in a very large excess of the reagent; insoluble in NH_4CO_2, and solution of NH_4Cl.
	Nothing .	in very dilute solutions.
Glucina .	*White* . .	bulky precipitate, insoluble in an excess of the reagent, but soluble in carbonate of ammonia.
Thorina . .	*White* . .	gelatinous precipitate, insoluble in an excess of the reagent
Yttria . . .	*White* . .	bulky precipitate (YO,HO); insoluble in an excess of the reagent.
Protoxyde of Cerium	*White* .	bulky precipitate (CeO,HO); insoluble in an excess of the reagent.
Zirconia . .	*White* . .	bulky precipitate (Zr_2O_3,HO); insoluble in an excess of the reagent.
Protoxyde of Manganese .	*White* . .	precipitate, in neutral solutions (MnO, HO), which by exposure turns
	Brown, . .	and then
	Black.	
	Nothing, .	if ammoniacal salts be present.

Sesquioxyde of Manganese . .	*Brown* . .	bulky precipitate.
Oxyde of Zinc . .	*White* . .	gelatinous precipitate (ZnO,HO); soluble in an excess of the reagent.
Protoxyde of Cobalt	. . . . .	With a small quantity of the reagent
	Blue	precipitate, with a larger quantity and exposure,
	Green . .	precipitate. Both precipitates are soluble in a large excess of the reagent, the liquor is *red*, and gradually becomes *brown* by exposure If enough NH_4Cl be present,
	Nothing . .	(See Table XXXI., Observation *a*.)
Protoxyde of Nickel	. . . . .	With a very small quantity of reagent,
	Green . .	precipitate, or turbidness, which disappears with the slightest excess, and the liquor has then a
	Fine blue colour . .	in which KO produces a *green precipitate*. (See Table XXXI, Observation *a*)
Protoxyde of Iron .	. . . . .	If the solution contains much NH_4Cl,
	Nothing, . .	but by exposure, or if no NH_4Cl be present, then a
	White . .	flocculent precipitate is produced, which, by exposure, soon becomes
	Grey . . . *Green* . . *Brown* . .	at the surface.
Peroxyde of Iron .		*Reddish-brown* precipitate, insoluble in an excess of the reagent.
Oxyde of Cadmium	*White* . .	precipitate (CdO,HO), in neutral solutions; soluble in a slight excess of the reagent.
Protoxyde of Lead .	*White* . .	precipitate, insoluble in an excess of the reagent. N.B. Solution of acetate of lead is not precipitated by NH , or only after some time.

Oxyde of Bismuth	*White* . .	precipitate, insoluble in an excess of the reagent.
Protoxyde of Uranium.	*Brownish black* . .	precipitate (UO,HO); insoluble in an excess of the reagent.
Peroxyde of Uranium .	*Yellow* . .	precipitate (uraniate of ammonia), insoluble in an excess of the reagent
Suboxyde of Copper	. . . . *Greenish* . . *Light blue* . *Dark blue* .	A small quantity produces a precipitate, becoming of a and then colour, by exposure. (See Table XVII., Observation *d.*)
Protoxyde of Copper	. . . . *Greenish* *Light blue* . *Dark blue* .	A small quantity produces a precipitate, becoming of a colour, and then by exposure (See Table XVII., Observation *d.*)
	Dark blue .	An excess of NH_3 produces a beautiful liquor, but no precipitate.
Oxyde of Silver .	*Light brown*	precipitate, soluble in the slightest excess of the reagent.
Suboxyde of Mercury .	*Black* . .	precipitate, insoluble in excess of the reagent
Peroxyde of Mercury .	*White* . .	precipitate, insoluble in excess of the reagent. (See Table IX, Observation *e*)
Protoxyde of Platinum .	*Green* . .	crystalline precipitate
Peroxyde of Platinum . .	*Yellow* . .	precipitate, soluble in a great excess of HO, soluble, with heat, in a great excess of NH_3
Oxyde of Palladium	. . . . . . *Colourless* .	An excess of NH_3 renders the solution (See Table XVII, Observation *g*)
Peroxyde of Rhodium	. . . . . *Yellow* . .	After a time, abundant precipitate (rhodate of ammonia); soluble in dilute HCl.

Binoxyde of Iridium .	. . . . . .	At first, the solution is
	Decolorised,	but by *boiling,* until the NH_3 is nearly expelled, the solution becomes
	Blue.	
Binoxyde of Osmium.	*Nothing.* .	at first, but afterwards the liquor becomes
	Brown . .	and a
	Brown . .	precipitate is produced.
Peroxyde of Gold . .	*Yellow* . .	precipitate.
Protoxyde of Tin .	*White* . .	precipitate (SnO,HO), soluble in an excess of the reagent.
Peroxyde of Tin .	*White* .	bulky precipitate (SnO,HO); soluble in a great excess of the reagent.
Oxyde of Antimony	*White* . .	abundant precipitate; settles after some time.
		N B. If tartaric acid is present, no precipitate is produced at first.
Protoxyde of Molybdenum .	*Brownish black* . .	precipitate (MoO,HO), insoluble in an excess of the reagent.
Binoxyde of Vanadium. .	*Greyish-white*	precipitate, soluble in an excess of the reagent, the solution has a *brown colour.* A large excess of NH_3 produces a
	Brown . .	precipitate, soluble in water.
Oxyde of Chromium .	*Bluish-grey*	precipitate (Cr_2O_3,HO), with a *tinge of violet:* the superincumbent liquor, by boiling, yields a precipitate.
Tellurous acid . .	*White* . .	abundant precipitate; soluble in an excess of the reagent.
Titanic acid .	*White* . .	bulky precipitate (TiO_2,HO), insoluble in an excess of the reagent.
Antimonic acid . .	*White* . .	precipitate; soluble in an excess of the reagent;
	Nothing. .	if organic matter is present.
Antimonious Acid	*White* . . .	precipitate; soluble in an excess of the reagent.

Permanganic acid	*Brown* . .	precipitate (MnO_2). Nitrogen is evolved.
Phosphates, Arseniates, Borates, Oxalates	of earths, *White* . .	precipitate. (See Table VI., Observations *a*, *i*; Table XXI., Observation *b*.)
Gallic acid .	*Reddish-brown* .	colour, becoming *green* or *purple*.

AMMONIO-NITRATE OF SILVER.

($AgO,NO_5 + 2NH_3$.)

Ammonio-nitrate of silver is prepared every time it is wanted by pouring carefully very dilute caustic ammonia into a solution of nitrate of silver, so that the light brown precipitate at first produced is redissolved by shaking. It is employed for the detection of arsenic and of arsenious acids in solutions which contain a free acid; but it is better, as we said in the tests for substances under the head Arsenious Acid, to neutralise the free acid with ammonia, and then to add a few drops of nitrate of silver; or else to add the nitrate of silver first, and then cautiously to drop some ammonia on the sides of the test-tube containing the liquor to be examined, so that it may float on the top of that liquor; if arsenious acid be present a ring of a yellow colour (AgO,AsO_3) will appear at the point of contact of the ammonia with the liquid; if arsenic acid is present, the ring has a reddish-brown colour.

AMMONIA PHOSPHATE OF SODA. (See Phosphate of Soda and Ammonia.)

AMMONIA SULPHATE OF COPPER.

This reagent is also prepared every time it is wanted by pouring into a solution of sulphate of copper as much dilute caustic ammonia as is just sufficient to redissolve the precipi-

tate at first formed. It is used then as a test for arsenious acid, with which it forms a yellowish-green precipitate; and for arsenic acid with which it forms a greenish-blue precipitate; both precipitates are immediately soluble in an excess of acid, or of ammonia.

ANTIMONIATE OF POTASH.

($KO,Sb_2O_5,7HO.$)

Antimoniate of potash, also called bimeta-antimoniate, or granular antimoniate of potash, is prepared by heating in a hessian crucible 4 parts of metallic antimony with 9 parts of nitre; the best way is to make the mixture of antimony and nitre, to project it by small portions at a time into the red-hot crucible, and to keep the mass for about half an hour in a pasty state, stirring it occasionally. The mass, after having cooled, is pulverised, washed with tepid water, and then boiled for a short time with water to remove the undecomposed nitrate and nitrite of potash; the washed residue is then dried, mixed with about two-fifths of its weight of carbonate of potash, and the mixture is exposed for half an hour to a bright red heat in a hessian crucible. The mass after it has cooled should be kept in a well-closed bottle. For use, pulverise a little of the compound, and dissolve 1 part of the pulverised product in about 20 parts of lukewarm water; when cold, filter.

Antimoniate of potash is sometimes used as a test for soda, which it throws down after some time in the state of a crystalline precipitate slightly soluble in water. This reagent is not much to be depended upon, and it cannot be used as a test for soda except in solutions which contain simply potash and soda, for it forms insoluble compounds with the alkaline earths and with most metallic oxydes.

AQUA REGIA.

This acid, called also nitro-chloric, or nitro-muriatic acid, or chloride of nitryle, is prepared by mixing about three measures of pure HCl with one of pure nitric acid. This mixture gives rise to the formation of a new compound, NO_3,Cl_2, by the reaction of the elements of the two acids on each other, thus:

$$2\,HCl + NO_5 = \underbrace{NO_3Cl_2}_{\text{Chloro-nitric acid}} + 2HO,$$

According to some chemists, this mixture consists of hyponitric acid and chlorine, the reaction being as follows :—

$$NO_5 + HCl = NO_4 + Cl + HO.$$

According to M. Baudrimont, however, the symbol for aqua regia should be NO_3Cl_2, that is to say, nitric acid, in which 2 equivalents of oxygen are replaced by 2 equivalents of chlorine (chloronitric acid). Gay Lussac considered aqua regia as being a hyponitric acid, in which 2 equivalents of oxygen are replaced by 2 equivalents of chlorine, namely, hypochloronitric acid (NO_2Cl_2), mixed with a nitrous acid, in which 1 equivalent of oxygen is replaced by 1 equivalent of chlorine, namely, chloronitrous acid (NO_2Cl), or in other words, a mixture of hypochloronitric acid, free chlorine and water, thus :—

$$NO_5,HO + 3HCl = NCl_2O_2 + Cl + 4HO.$$

According to Gerhardt, nitric and hydrochloric acid, when mixed together, produce, by double decomposition, water and chloride of nitryle (NO_4Cl), thus :—

$$\left.\begin{matrix} NO_4O \\ HO \end{matrix}\right\} + \left.\begin{matrix} H \\ Cl \end{matrix}\right\} = \left.\begin{matrix} HO \\ HO \end{matrix}\right\} + \begin{matrix} NO_4 \\ Cl. \end{matrix}$$

Howbeit, aqua regia is a liquid of a yellowish-red colour, which has the property of dissolving gold, platinum, and sulphuret of mercury, which are insoluble in both hydrochloric and nitric acid separately, and peroxyde of tin and of antimony, which are insoluble in nitric acid. All the substances which are dissolved by aqua regia are always converted by it into chlorides.

A mixture of chlorate of potash (KO,ClO_5), and of dilute HCl, yields also a liquid analogous to aqua regia, and which can dissolve metals and sulphurets like that acid.

Aqua regia dissolves not only all the substances which are soluble in either nitric or hydrochloric acid, but several others on which these acids separately have not the slightest action. (See Table I.—E, Observations *n*, *p*, *r*, *s*, *u*.)

ARSENIOUS ACID.

(AsO_3.)

Arsenious acid is only used as a test, in conjunction with caustic potash, for the detection of acetic acid and acetates. (See Table I.—C, col. 19.)

ARSENIATE OF SODA.

Arseniate of soda is sometimes used as a test for salts of barium, with which it forms a white precipitate, insoluble in HO, soluble in NO_5.

BARYTA (CAUSTIC).

(BaO.)

(See also Hydrate of Baryta, and Baryta Water, BaO+aq.)

Aqueous solution of caustic barytes is prepared by calcining,

in a porcelain crucible or retort, nitrate of barytes at a white heat, or, in fact, until ruddy fumes or oxygen are no longer evolved. The residue consists of caustic barytes, which may then be dissolved in water, and kept for use in well-stoppered bottles. This method, however, is rather inconvenient, because the nitrate of barytes swells considerably and necessitates the use of very large vessels to prevent overflowing.

A better method is the following :—mix intimately together 8 parts of ordinary or of natural sulphate of barytes, previously reduced into very fine powder, with 1 part of charcoal, also in very fine powder, and 2 parts of flour or of rosin, and expose the mixture to a white heat for about one hour in a hessian crucible. The residue having become perfectly cold, boil it with water and some black oxyde of copper until a portion of the liquid being filtered, fails to produce a black precipitate by testing with $PbO,\bar{A}$. Throw the whole mass in a filter whilst boiling hot, dilute the filtrate with water, and keep it for use in glass-stoppered bottles. If on cooling crystals are deposited, they consist of BaO,HO, and should be dissolved by adding more water.

Solution of BaO in water, or baryta-water, as it is called, is principally used for the detection of carbonic acid, for which purpose, however, a solution of sulphuret of barium, or lime-water, answers just as well. It is also used for precipitating magnesia from solutions which contain alkalies; but it precipitates, in fact, all the alkaline earths from the solutions of their salts, and it may be substituted to the soluble salts of barium for the detection of sulphuric acid.

In the solid state hydrate of caustic baryta (BaO,HO), (which see), is used as a flux for certain silicates.

BORAX.

BIBORATE OF SODA ($NaO,2BO_3,10HO$), AND NEUTRAL BORATE OF SODA ($NaO,BO_3,8HO$).

Borax is generally found in commerce in a state of sufficient purity for analytical purposes, yet it is advisable to recrystallise it. It occasionally contains a small quantity of organic matter, owing to which, the bead produced by fusion has a greyish colour; by continuing the heat, however, the greyish colour disappears, and a clear bead is obtained.

The neutral borate of soda precipitates a solution of *perchloride of iron*, but on applying heat the precipitate is redissolved, and the liquor has then a dark red colour, resembling that which *acetic acid* produces in such a solution, and hence the operator might be led into error in attributing the reaction to the presence of the latter acid.

To be fit for blowpipe experiments, a portion of the salt dissolved in water and acidified by NO_5, should not be rendered turbid by either *nitrate of silver* or *nitrate of barytes*. Neither should solution of *carbonate of soda* or of *potash* produce anything in aqueous solution of borax. If the solution is affected by either of these reagents, the borax must be recrystallised.

For use, the pure crystals of borax should be gently heated in a platinum crucible until they cease to swell, the fused mass is then reduced to powder, and kept in a bottle.

Borax is one of the most useful of the reagents for blowpipe operations, a platinum wire bent into a hook or loop at one of its extremities being generally employed as a support of the substance to be fused with it,—because upon charcoal it is impossible to obtain a well-formed bead, the

borax spreading upon the charcoal and agglomerating into a round bead only with difficulty, and after protracted blowing. The method of using the salt is as follows:—the platinum-wire hook is first moistened with a little saliva, and then plunged into the pulverised borax; the adhering mass is first fused in the oxydising or outer flame into a clear bead, and a small portion of the compound to be thus examined (and which need not be more than the tenth or even the hundredth part of the weight of the borax) is added to the fused bead by touching the pulverised compound with the fused borax whilst red hot, or after moistening it with saliva if it be cold; the charged hook is then exposed to the action of both the inner and outer flames of the blowpipe, the result being either a colourless, or a coloured bead, according to the nature of the oxyde under examination.

COLOURS IMPARTED TO A BEAD OF BORAX BY METALLIC OXYDES IN BOTH FLAMES OF THE BLOWPIPE.

Names of Metallic Oxydes	Colourless in both Flames	
Potash . . .	Glucina . . .	Thorina.
Soda . . .	Yttria . . .	Silica.
Lithia . . .	Zirconia . . .	Tin, only slightly soluble in borax.
Baryta . . .	Tantalic acid.	
Strontia . . .	Zinc . . .	Silver, becomes turbid, if in abundance, by intermittent blowing.
Lime .	Cadmium.	
Magnesia . .	Alumina.	

Names of Metallic Oxydes	Colours of the Bead in the Oxydising or Outer Flame.	Colours of the Bead in the Deoxydising or Inner Flame
Molybdenum .	Colourless . .	Brown, or reddish-brown
Tungsten . . .	" . . .	Yellow, when hot; brown, when cold.
Tellurium .	" . . .	Grey ; opaque.
Bismuth . .	" . . .	Grey; opaque.
Titanium . . .	" . . .	Violet, or bluish-black, if abundant.
Antimony . .	Colourless, when cold, yellow, when hot .	Grey, opaque.
Lead . . .	Colourless, or almost so, when cold, yellow, when hot . .	Metallic lead.
Chrome . . .	Green . . .	Fine green.
Copper (protoxyde)	Green	Reddish-brown.
Cerium . .	Red, while hot, then yellow, colourless, when cold .	Colourless.
Uranium . . .	Yellow . . .	Dingy green.
Vanadium . .	Yellow . .	
Nickel . . .	Red, while hot, colourless, when cold .	Grey ; opaque.
Iron (peroxyde) .	Reddish, while hot, then yellow, and, finally, colourless, when cold . .	Bottle-green.
Manganese . .	Violet, or amethyst . .	Colourless.
Cobalt . . .	Blue . . .	Blue.

Colours of the Bead in the Oxydising or Outer Flame		Colours of the Bead in the Deoxydising or Inner Flame.	
Colourless	Baryta.	Colourless	Baryta.
,,	Strontia	,,	Strontia
,,	Lime.	,,	Lime.
,,	Magnesia.	,,	Magnesia.
,,	Glucina.	,,	Glucina.
,,	Yttria	,,	Yttria.
,,	Zirconia.	,,	Zirconia.
,,	Tantalic acid.	,,	Tantalic acid.
,,	Titanic acid.	,,	Zinc.
,,	Zinc.	,,	Cadmium.
,,	Cadmium.	,,	Alumina.
,,	Silver.	,,	Thorina.
,,	Alumina.	,,	Silica.
,,	Thorina.	,,	Tin
,,	Silica.		
,,	Tellurium.		[All the above oxydes yield a light-coloured glass in the exterior flame]
,,	Bismuth.		
,,	Antimony.		
,,	Tungsten.		
,,	Molybdenum.	,,	Cerium.
,,	Tin.	,,	Manganese.
Green	Chromium.	Green	Chromium.
,,	Copper (protoxyde)	,,	Vanadium.
Yellow	Vanadium.	,,	Uranium.

Colours of the Bead in the Oxydising or Outer Flame		Colours of the Bead in the Deoxydising or Inner Flame.	
Yellow	Uranium.	Green	Iron.
„	Lead; almost colourless, in cooling.	Yellow	Tungsten.
		Brown, or reddish-brown.	Molybdenum.
Red	Cerium.	„	Copper (protoxyde)
„	Iron (peroxyde).	Blue	Cobalt.
„	Nickel.	Violet	Titanium.
	[All these become colourless, or almost so, in cooling.]	Grey	Antimony.
		„	Tellurium.
		„	Nickel.
Blue	Cobalt.	„	Bismuth.
Violet	Manganese.	„	Silver.

BICARBONATE OF POTASH ($KO,2CO_2$) OR OF SODA ($NaO,2CO_2$).

Bicarbonate of potash and bicarbonate of soda behave with substances in the same manner, and they are found in commerce, especially the last, in a state of almost absolute purity, a trace of sulphate only being present.

These bicarbonates are much less soluble than their simple carbonates, and the solution is nearly neutral to test papers. Carbonic acid is disengaged when their solution is made to boil; the solution, in fact, begins to disengage CO_2 when heated at about 160° Fahr., and is converted into neutral carbonate by boiling. The same thing happens with bicarbonate of soda.

When bicarbonate of potash contains any carbonate, it deliquesces by exposure.

The solution of bicarbonate of potash or of soda should not precipitate sulphate of magnesia; in fact, this is a way to distinguish bicarbonate of potash or of soda from their neutral carbonates, since the first do not precipitate the salts of magnesia, whereas the second do so.

For use in solution, dissolve 1 part of the salt in about 8 or 10 parts of water.

Bicarbonate of potash or of soda are principally used to distinguish magnesia from alumina and baryta, strontia and lime from protoxyde of manganese.

REACTIONS.

Potash . . .	*Nothing.*	
Soda . . .	*Nothing.*	
Lithia . . .	*Nothing.*	
Baryta . .	*White* .	precipitate; sparingly soluble in a great quantity of water
Strontia . .	*White* . .	precipitate, as with baryta.
Lime . . .	*White* . .	precipitate; as with baryta.
Magnesia . .	*Nothing* *White* . .	in the cold; but by boiling precipitate
Alumina .	*White* . .	precipitate (Al_2O_3,HO), insoluble in an excess of the reagent, an effervescence takes place at the same time.
Glucina .	*Bulky flocculent* .	precipitate (GlO,CO_2), soluble in a great excess of the reagent.
Yttria . . .	*White* . .	bulky precipitate (YO,CO_2); soluble in a great excess of the reagent.
Protoxyde of Cerium	*White* .	bulky precipitate (CeO,CO_2), soluble in a great excess of the reagent. (See Table XXXI., Observation *e.*)
Zirconia . .	*White* . .	bulky precipitate (Zr_2O_3,CO_2); soluble in a great excess of the reagent.

Protoxyde of Manganese	*White* . .	bulky precipitate. If the solution is dilute, or if ammoniacal salts are present, some time is required for the precipitate to appear.
Sesquioxyde of Manganese . .	*Brown* . .	voluminous precipitate (Mn_2O_3,HO).
Oxyde of Zinc . .	*White* . .	precipitate (ZnO,CO_2); carbonic acid is disengaged at the same time.
Protoxyde of Cobalt	*Pink*. . .	precipitate (CoO,CO_2).
Protoxyde of Nickel	*Light apple-green*.	precipitate (NiO,CO_2); accompanied by a disengagement of CO_2.
Protoxyde of Iron	*White* . .	precipitate (FeO,CO_2); accompanied by a disengagement of CO_2.
Peroxyde of Iron .	*Light reddish brown*	precipitate, CO_2 being at the same time disengaged, especially by boiling, and the precipitate then becomes *darker*.
Oxyde of Cadmium	*White* . .	precipitate (CdO,CO_2); in neutral solutions CO_2 is disengaged at the same time.
Protoxyde of Lead .	*White* . .	precipitate (PbO,CO_2); disengagement of CO_2.
Oxyde of Bismuth	*White* . .	precipitate (BiO,CO_2); disengagement of CO_2.
Protoxyde of Uranium . .	*Dingy greenish*	precipitate (UO,CO_2); soluble in a great excess of the reagent.
Suboxyde of Copper . .	*Yellow* . .	precipitate (Cu_2O,CO_2).
Protoxyde of Copper . .	*Light greenish-white*	precipitate; soluble in a great excess of the reagent.
Oxyde of Silver .	*White* . .	precipitate (AgO,CO_2); soluble in NH_3.

Suboxyde of Mercury . .	*White* . .	precipitate (Hg_2O,CO_2); turning *black* by boiling, CO_2 is disengaged.
Peroxyde of Mercury . .	*Reddish-brown* .	precipitate. In solution of perchloride, *nothing* at first; after a time, *dark reddish-brown* precipitate.
Peroxyde of Platinum . .	*Yellow* . .	precipitate ($PtCl_2,KCl$); especially by adding HCl; insoluble in a great excess of the reagent, even with the help of heat.
Oxyde of Palladium .	*Brown* . .	precipitate; soluble in a great excess of the reagent.
Binoxyde of Iridium .	*Nothing*. .	at first; after a time, the solution is *decolorised.*
Binoxyde of Osmium .	*Nothing*. .	at first; by boiling, *black* precipitate.
Peroxyde of Gold .	*Nothing.*	
Protoxyde of Tin .	*White* . .	precipitate (SnO,HO); insoluble in a great excess of the reagent.
Peroxyde of Tin .	*White* . .	precipitate (SnO_2,HO); insoluble in a great excess of the reagent.
Oxyde of Antimony .	*White* . .	abundant precipitate. If *tartaric acid* is present, the precipitate takes *a long time* to appear.
Protoxyde of Molybdenum .	*Brownish-black* . .	precipitate (MoO,HO); slightly soluble in a great excess of the reagent.
Binoxyde of Vanadium .	*Grey* . . .	precipitate (VaO,HO); soluble in a great excess of the reagent.
Oxyde of Chromium . .	*Light greenish* . .	precipitate (Cr_2O_3); and the superincumbent liquor is *greenish.*

Tellurous acid	*White* . .	abundant precipitate; soluble in a great excess of the reagent.
Titanic acid .	*White* . .	bulky precipitate (TiO_2,HO); insoluble in a great excess of the reagent.
Antimonic acid . .	*White* . .	precipitate; partly soluble in a great excess of the reagent.
Antimonious acid .	*White* . .	precipitate; partly soluble in a great excess of the reagent. If *organic matter* is present, *nothing*.
Zirconia . .	*White* . .	precipitate; soluble in a great excess of the reagent. (See Table XXXI., Observation *a*.)
Thorina, Yttria .	*White* . .	precipitate; soluble in a great excess of the reagent. (See Table XXXI., Observation *e*.)

BICARBONATE OF SODA ($NaO,2CO_2$). (See Bicarbonate of Potash.)

BICHLORIDE OF MERCURY. (See Perchloride of Mercury.)

BICHLORIDE OF PLATINUM.

($PtCl_2$.)

Bichloride of platinum (perchloride of platinum) is easily prepared by dissolving in aqua regia, and at a gentle heat, pieces of platinum, previously cleaned by boiling them for a short time in nitric acid. The aqua regia should consist of 2 parts of concentrated HCl and 1 part of NO_5, a little of the latter acid being added from time to time until all the platinum has disappeared. The solution, which is of a brownish-red colour, is then evaporated to dryness by means of a steam bath, in order to expel the excess of acid, and 1 part of the dry residue is then dissolved in 8 or 10 parts of pure water.

Perchloride of platinum is used only for the detection of potash and of ammonia, in the concentrated solutions of which it forms a yellow precipitate, which is a double chloride of platinum and potassium or ammonium. It is also used as a test for SnO.

REACTIONS.

Potash . . .	*Pale yellow*	precipitate; sparingly soluble in water; insoluble in alcohol. In dilute solutions, a long time elapses before any precipitate makes its appearance, and it is then of a *reddish colour*, and crystalline. (See Table XX., Observations *d*, *e*.)
Lithia		In dilute alcoholic solutions,
	Very slight turbidness,	almost invisible; in concentrated solution the turbidness is more distinct.
Ammonia .	*Pale yellow*	precipitate, exactly as with potash.
Protoxyde of Tin	. . . . . .	In dilute solutions
	Reddish-brown .	precipitate.
		In concentrated solutions
	Blood-red colour .	(See Table IV., Observation *z*.)

BICHROMATE OF POTASH.

$(KO, 2CrO_3)$

Bichromate of potash is found in a state of considerable or even of perfect purity in commerce, and on that account it is often used in preference to the neutral chromate (KO, CrO_3) as a test. The reactions are the same as with the neutral chromate, to which, in certain cases, it is even preferable. (See Table VI., Observation *f*, and Table X., Observation *c*.)

It is used also in the solid state to distinguish strychnine from all other alkaloids. For this purpose a minute crystal

of bichromate is placed by the side of the strychnine, and a few drops of concentrated SO_3,HO are poured upon the two, which, on being stirred so as to touch each other, will then exhibit a beautiful purple colour. This test is well defined only when the strychnine is in a pure state.

For use in the wet way, dissolve 1 part of the bichromate in about 10 parts of water.

BIMETA ANTIMONIATE OF POTASH. (See Antimoniate of Potash.)

BINOXALATE OF POTASH.

(KO, $(2\bar{O})$, 3HO, or KO, $2C_2O_3$ + 3HO.)

The solution of the binoxalate of potash of commerce (sorrel salt) may generally be used instead of oxalic acid. It is necessary, however, to bear in mind that the commercial salt is often adulterated with *tartrate of potash*, and is contaminated besides by organic impurities. The easiest method of testing its purity consists in boiling a portion of it with some concentrated SO_3,HO, this decomposes the oxalate, an evolution of gases takes place, and if the liquor left is colourless, the salt is pure, but if it turns brown or black, and sulphurous acid is evolved with the other gases by protracted boiling, then *tartaric acid, or some other organic substance,* is present. *Bisulphate of potash* is detected by dissolving a portion of the salt, adding a little HCl, and testing the solution with chloride of barium, which will then produce a white precipitate of sulphate of baryta.

REACTIONS.

Baryta . . .	*Nothing* . .	even in concentrated solutions: but if NH_3 be added, an abundant *white* precipitate is produced, soluble in acids.

Strontia . .	*Slight turbidness*	at first; but by adding NH, *white precipitate.*
Lime . . .	*White* . .	precipitate ($CaO,\bar{O}$); augmenting by repose, *and especially by adding* NH_3.
Magnesia . .	*Nothing.*	
Glucina . .	*Nothing.*	
Protoxyde of Iron .	*Yellow* . .	precipitate ($FeO,\bar{O}$); soluble in HCl.

BORAX. (See Biborate of Soda.)

BISULPHATE OF POTASH.

($KO,2SO_3,HO$.)

Bisulphate of potash is a very acid salt, which easily fuses when heated, and becomes then very fluid. Heated to a bright red heat, it is decomposed into sulphurous acid, mixed with sulphuric acid. This property of bisulphate of potash renders it available in the inorganic analysis of certain substances which cannot be attacked by heating them with concentrated sulphuric acid, because that acid boils and evaporates at 620° Fahr., whereas, by fusion with bisulphate of potash, they are readily decomposed.

Remember that bisulphate of potash should not be fused in platinum crucibles. (See Preliminary Observations, page vii.)

Bisulphate of potash is found in commerce sufficiently pure for analytical purposes, but it may be easily prepared by mixing the neutral sulphate of potash, previously pulverised, with half its weight of concentrated sulphuric acid, and evaporating the whole to perfect dryness in a platinum capsule placed under the hood of a chimney. The calcination should be continued until fumes of sulphuric acid are no longer evolved. The calcined mass is then dissolved in hot water and

left to crystallise. The crystals so obtained are in the shape of colourless prisms, much more soluble than those of the neutral sulphate, and require only twice their weight of cold water, or their own weight of boiling water for the purpose.

This salt is chiefly used for the detection of boracic acid in the borates, and of bromine and iodine in the combinations of these substances. Thus, when borates are triturated with about 3 or 4 times their weight of a mixture of 1 part of fluorspar and 4½ parts of bisulphate of potash and a portion of the mass, *slightly moistened*, is held at the point of the inner flame of the blowpipe on the hook of a platinum wire, a fine green flame is at once produced, but it lasts only an instant.

The use of that salt as a test for bromine and iodine has been described in Table I., A, cols. 4 and 6.

BLACK FLUX. (See Flux.)

BLUE LITMUS PAPER. (See Test-papers.)

BITARTRATE OF POTASH ($KO,HO,\bar{T}$)—CREAM OF TARTAR.

The bitartrate of potash (cream of tartar) of commerce is sufficiently pure for analytical purposes. It is used as a test for distinguishing oxyde of antimony from oxyde of bismuth, the first being soluble in a concentrated solution of this salt, whilst the second is insoluble therein; and for the preparation of the black and of the white fluxes. (See Flux.)

BORACIC ACID.

(BO_3.)

Boracic acid may be easily obtained by decomposing borax with a slight excess of sulphuric acid. The crystalline mass

which falls down is boracic acid, which should be purified by washing it with *cold* water until the filtrate is no longer rendered turbid when acidified with nitric acid and tested with BaCl.

Boracic acid is a special reagent for the blowpipe, and is seldom used. Berzelius has recommended it, in conjunction with iron, as a means of detecting phosphoric acid; the process, however, is insufficient for detecting small quantities of that acid, and it is inapplicable if the compound under examination contains either sulphuric or arsenic acids, or any metallic oxyde capable of being reduced by iron. The process, however, consists in fusing a portion of the substance with boracic acid upon charcoal, forcing a small piece of fine steel wire through the bead when in fusion, and then heating the whole again strongly in the inner flame of the blowpipe. After cooling, the bead is struck gently with a hammer upon an anvil so as to detach the phosphuret of iron, in the form of a round metallic grain, which is magnetic, and flies to pieces when struck with a smart blow.

This process is far from being satisfactory or conclusive.

BORATE OF SODA (NEUTRAL). (See Biborate of Soda.)

BORAX. (See Biborate of Soda.)

BRAZIL WOOD PAPER. (See Test Papers.)

BROWN OXYDE OF LEAD.

(PbO_2)

Brown oxyde of lead, called also *peroxyde* or *puce oxyde of lead,* is a metallic acid (plumbic acid), of a brown or dark chocolate colour, insoluble in water. It is an energetic

oxydiser. If a little of it be dropped into ammonia, it absorbs it, and is thereby partially reduced, water and nitrate of ammonia being produced.

Plumbic acid is easily prepared as follows:—Reduce red lead (plumbate of protoxyde of lead—PbO_2PbO) to powder, mix it in a Florence flask with an excess of nitric acid, diluted with 2 or 3 times its bulk of water, and heat the mixture to the boiling point, stirring all the while. The brown insoluble powder left should be separated by filtering, and washed until the filtrate ceases to have an acid reaction upon blue litmus paper; it is then dried at 212° by means of a steam-bath, and the dry powder is kept in a bottle for use.

Another easy method of preparing brown oxyde of lead, is the following:—Dissolve, in a suitable vessel, 5 ounces of acetate of lead in cold water, and in another vessel dissolve 4 ounces of crystallised carbonate of soda also in cold water; mix the two solutions together. This of course will produce a thickish mass, due to a precipitation of the lead in the state of carbonate; add as much water as is requisite to bring it to the consistence of cream, and pass a stream of chlorine through the creamy mass until the whole is converted into a deep brown precipitate. Collect the precipitate on a filter, and wash it thoroughly. It is the brown oxyde in question.

Brown oxyde of lead is used in chemical analysis for the purpose of separating sulphurous acid from other gases, sulphurous acid being absorbed by it with great rapidity, and sometimes with great elevation of temperature, sulphate of lead being thereby produced. Mixed with water containing sulphurous acid, and agitated therewith, the sulphurous acid is absorbed, and sulphate of lead is also produced—thus:

$$PbO_2 + SO_2 = PbO,SO_3.$$

In conjunction with peroxyde of manganese and concentrated sulphuric acid it serves to detect the presence of chlorides, because when a compound containing a chloride is

mixed with peroyxde of manganese and brown oxyde of lead, if concentrated sulphuric acid be poured upon the mixture and heat applied, chlorine is then evolved, recognisable by its odour and colour, and also by its property of bleaching moist litmus paper. .

If it be mixed with solutions of oxalic acid and of acid oxalates, carbonic acid is disengaged.

Mixed with a solution of caustic potash, it serves to transform sesquioxyde of chromium into yellow chromate of lead. (See Table XVIII., Observation *h.*)

It is also used for the detection of the most minute traces of manganese. The compound is first dissolved in an appropriate menstruum, PbO_2 and some dilute NO_5 are then added, and the whole being boiled assumes a fine purple colour, if the compound under examination contained the smallest trace of manganese. The purple colour is due to the formation of permanganic acid (Mn_2O_7)—thus:

$$2MnO + 3PbO_2 + NO_5 = Mn_2O_7 + PbO, NO_5.$$

BRUCINE.

($C_{44}H_{25}N_2O_7$.)

Brucine is one of the alkaloids obtained from nux vomica, and as a test it is only used for the detection of nitric acid. The modus operandi is as follows:—Pour into a glass vessel 50 or 60 grains measure of concentrated sulphuric acid (of course perfectly free from nitric acid), and then a few drops of the liquor in which the presence of nitric acid is suspected; stir the whole with a glass rod, and with the moist extremity of the glass rod introduce a small particle of brucine; stir the whole well; if any nitric acid is present, the liquid becomes first *deep red,* then *yellowish-red,* and finally quite *yellow, especially on applying heat.* If now a little solution of *proto-*

chloride of tin ($SnCl$) be poured into the yellow liquor, it becomes *purple*. If no nitric acid is present the brucine does not even dissolve. According to Mr. Berthemot $\frac{1}{10000}$ of nitric acid may thus be detected. The operator should ascertain that the sulphuric acid employed contains no nitric acid, by testing a portion of it separately with brucine in another tube.

CARBAZOTIC ACID. (See Picric Acid.)

CALOMEL. (See Subchloride of Mercury.)

CARBONATE OF AMMONIA.

(NH_4O,CO_2) or $(2NH_4O),3CO_2+2HO$.

Carbonate of ammonia, or rather, sesquicarbonate of ammonia, is largely prepared for pharmaceutical and other purposes by subliming a mixture of sal-ammoniac and chalk. It is sometimes called neutral carbonate, but it contains, in reality, less base than is required to form that salt, for even when freshly prepared, its composition is very near that of a sesquicarbonate $(2NH_4O,)3CO_2$. The neutral carbonate is in reality unknown, except in solution in water, or in alcohol.

This salt is generally pure, but it is sometimes contaminated by NH_4Cl,—NH_4O,SO_3,—by organic substances,—and occasionally by PbO,CO_2, or by salts of lime: such a carbonate of ammonia is, of course, unfit for analytical purposes, but it is, as we said, generally found in commerce sufficiently, and often quite pure.

Pure carbonate of ammonia should evaporate without residue when exposed to a red heat in a platinum capsule, and its

aqueous solution previously supersaturated with NO_5 should not be discolored, nor rendered turbid, nor precipitated by

BaO,NO_5, which otherwise would indicate the presence of		Sulphates.
AgO,NO_5,	ditto	Chlorides.
HS,	ditto	Metallic oxydes.

For use one part of the transparent or superficially effloresced crystals should be scraped or washed clean, and then dissolved in about four parts of water, after which one part of caustic ammonia is further added to the whole in order to convert it into the neutral carbonate.

Carbonate of ammonia has nearly the same reactions as carbonate of potash, though there are several bases which the latter precipitates, and on which carbonate of ammonia has no action, especially when ammoniacal salts or free acids are present, because a soluble double salt of ammonia and of the base is produced. It is principally used for separating baryta, strontia, and lime from magnesia, the latter substance not being precipitated by NH_4O,CO_2 in the presence of ammoniacal salts. Carbonate of ammonia is in a great many cases preferable to the carbonates of the fixed alkalies, because the ammoniacal salts which result from its use can always be evaporated or expelled by heat, which is not the case, of course, when the fixed carbonates are employed.

REACTIONS.

Baryta . .	*White* . .	precipitate (BaO,CO_2). (See Table VI., Observation *b*.)
Strontia . .	*White* . .	precipitate (StO,CO_2). (See Table VI., Observation *b*.)
Lime . . .	*White* . .	precipitate (CaO,CO_2). (See Table VI., Observation *b*.)
Magnesia . .	*Nothing* . .	in the cold;
	White . .	precipitate, by boiling. (See Table VI., Observation *b*.)

Alumina . .	*White* . .	precipitate (Al_2O_3,HO); insoluble in an excess of the reagent, and in water, an effervescence is produced at the same time.
Glucina . .	*White* . .	flocculent bulky precipitate (GlO,CO_2); soluble in a great excess of the reagent, but reappears by boiling.
Thorina . .	*White* . .	precipitate; soluble in an excess of the reagent.
Yttria . .	*White* . .	bulky precipitate (YO,CO_2); soluble in a very large excess of the reagent.
Protoxyde of Cerium . .	*White* .	bulky precipitate (CeO,CO_2); soluble in an excess of the reagent.
Zirconia . .	*White* .	bulky precipitate (Zr_2O_3,CO_2); soluble in an excess of the reagent.
Protoxyde of Manganese .	*White* . .	precipitate (MnO,CO_2).
Sesquioxyde of Manganese . .	*Brown* . .	voluminous precipitate (Mn_2O_3,HO).
Oxyde of Zinc . .	*White* . .	precipitate (ZnO,CO_2); soluble in an excess of the reagent.
Protoxyde of Cobalt	*Red* . . .	precipitate (CoO,CO_2), (in neutral solutions); soluble in NH_4Cl.
Protoxyde of Nickel	*Apple-green*	precipitate (NiO,CO_2); (in neutral solutions) soluble in an excess of the reagent; the liquor is *greenish blue.*
	Nothing . .	in acid solutions.
Protoxyde of Iron .	*White* . .	precipitate (FeO,HO); becoming
	Grey . . .	then
	Green . .	and finally
	Brown . .	by exposure.
Peroxyde of Iron .	*Reddish-brown* .	precipitate (Fe_2O_3,HO).

Oxyde of Cadmium	*White* . .	precipitate (CdO,CO_2); insoluble in an excess of the reagent. (See Table XVII., Observation *e.*)
Protoxyde of Lead .	*White* . .	precipitate (PbO,CO_2); insoluble in an excess of the reagent.
Oxyde of Bismuth .	*White* . .	precipitate (BiO,CO_2); immediately, or after some time, especially in a warm place.
Protoxyde of Uranium .	*Dingy greenish*	precipitate (UO,CO_2); soluble in an excess of the reagent.
Peroxyde of Uranium .	*Yellow* . .	precipitate (U_2O_3,CO_2); soluble in an excess of the reagent.
Suboxyde of Copper . .	*Light-blue* .	colour, becoming *dark-blue* by exposure; CO_2 is disengaged.
Protoxyde of Copper . .	*Slightly greenish*	precipitate, with small quantity of reagent; soluble in an excess thereof; the liquor is then deep *blue.*
Oxyde of Silver .	*White* . .	precipitate (AgO,CO_2); soluble in an excess of the reagent.
Suboxyde of Mercury . .	*Grey* . . .	precipitate, with a small quantity of the reagent.
	Black . .	precipitate, with a large quantity of the reagent.
Peroxyde of Mercury . .	*White* . .	precipitate.
Protoxyde of Platinum .	*Nothing.*	
Peroxyde of Platinum .	*Yellow* . .	precipitate, especially by adding HCl.
Peroxyde of Palladium .	*Nothing* . .	but an excess of reagent renders the solution *colourless.*

Peroxyde of Rhodium	*Nothing* . . *Yellow* . .	at first; precipitate, after a long time.
Binoxyde of Iridium.	*Nothing*. .	at first; after a time the solution is *decolourised.*
Binoxide of Osmium	*Nothing*. . *Brown* . . *Brown* . .	at first; after a time the solution becomes and finally a precipitate is produced.
Peroxyde of Gold .	*Yellow* . .	precipitate; (in neutral solutions) CO_2 is disengaged.
Protoxyde of Tin . .	*White* . .	precipitate (SnO,HO); insoluble in excess.
Peroxyde of Tin .	*White* . .	precipitate (SnO_2HO); insoluble in an excess of the reagent.
Oxyde of Antimony .	*White* .	bulky precipitate. If *tartaric acid* is present, *nothing.*
Protoxyde of Molybdenum .	*Brownish-black* . .	precipitate (MoO,HO); soluble in an excess of the reagent.
Binoxyde of Molybdenum .	*Light-brown*	precipitate; soluble in an excess of the reagent.
Binoxyde of Vanadium .	*Greyish-white* . .	precipitate (VaO_2,HO); the solution has a *brown colour.*
Oxyde of Chromium .	*Light-green*	precipitate (Cr_2O_3,HO); the superincumbent liquor is *greenish.*
Tellurous acid . .	*White* . .	precipitate; soluble in an excess of the reagent.
Titanic acid	*White* .	bulky precipitate (TiO_2,HO), insoluble in an excess of the reagent.
Permanganic acid .	*Brown* . .	precipitate (MnO_2).

CARBONATE OF BARYTES.

(BaO,CO_2.)

Carbonate of barytes is but seldom found in commerce in a state of sufficient purity; it is generally contaminated by fixed alkalies and other salts, probably from careless or insufficient washing.

Carbonate of barytes may be easily prepared by mixing a solution of chloride of barium with an excess of carbonate of ammonia; the precipitate produced is carbonate of barytes, which should be thoroughly washed with boiling water.

Whether carbonate of barytes contains any soluble salts is of course easily ascertained by boiling a certain portion of it in water and filtering; the filtrate should not leave any appreciable residue by evaporation. The other impurities are recognised by dissolving a portion of it in HCl, adding an excess of diluted sulphuric acid, and filtering; a few drops of the filtrate evaporated upon a slip of platinum foil should not leave the slightest residue.

Carbonate of barytes is used for precipitating several weak bases, such as Mn_2O_3—Al_2O_3—Cr_2O_3—Fe_2O_3 and separating them from other oxydes, such as, for example, from MnO —ZnO—CoO—NiO, for which purpose it is added in powder and in great excess to the solution. (See my edition of Rose's Practical Treatise of Chemical Analysis.)

CARBONATE OF LIME.

(CaO,CO_2.)

Chalk is used in the same way and for the same purposes as carbonate of barytes, but as it does not separate from the liquor to which it is added so rapidly as the latter carbonate

—as it yields, with hydrochloric acid, solutions of chloride of calcium which filter much more slowly than chloride of barium, and is much less easily and rapidly separated from the liquor than baryta, which an addition of SO_3 at once completely precipitates—carbonate of barytes is generally a preferable reagent for the purpose.

CARBONATE OF POTASH, OR OF SODA.

(KO,CO_2, or NaO,CO_2.)

The carbonate of potash, or of soda, of commerce is not sufficiently pure for analytical purposes.

It may be easily obtained by recrystallising the bicarbonate of these bases, igniting it in a clean iron pan, treating the ignited mass with boiling water, and filtering.

Carbonate of potash may also be prepared from the bitartrate of that base, ordinarily called cream of tartar ($KO, HO,\overline{T}$), as follows:—Reduce to powder a certain quantity of cream of tartar, if not already in that state, mix it with its own weight of water, acidified with a few drops of HCl, and digest it at the heat of a steam-bath for a while, taking care to stir frequently; throw the mass upon a linen filter placed in a funnel, and allow it to drain; wash it then thereon by drenching it with cold water until the filtrate, acidified with NO_5, is no longer precipitated by AgO,NO_5. This done, dry the mass in the filter, mix it with half its weight of pure KO,NO_5, and project it by small portions at a time in an iron pan (cast iron) kept at a dark red heat. Having thus deflagrated the whole mixture, increase the heat to a cherry red, pushing the portions which adhere to the sides into the rest of the mass, and continue the heat until, on dissolving in water a portion of the mass taken as sample and filtering, it produces a perfectly clear and colourless liquor. Treat then the whole mass in the pan, after it has

cooled, with boiling water, filter, and evaporate the filtrate in a porcelain, or better still, in a silver capsule, until a hard crust is formed on the surface. Remove the fire, stir the mass until cold, put the crystalline mass to drain, and keep it for use.

The principal impurities by which carbonate of potash and of soda may be contaminated are KO,SO_3, or NaO,SO_3,—KCl or $NaCl$, and in the case of carbonate of potash Al_2O_3 and SiO_3. These carbonates, to be fit for analytical purposes, should be perfectly white, and their aqueous solution must not be precipitated or rendered turbid, after supersaturation with NO_5, by

$BaCl$, which otherwise indicates the presence of		Sulphates.
AgO,NO_5,	ditto	Chlorides.
NH_4O,CO_2,	ditto	Alumina.

If, after having supersaturated the solution of KO,CO_2 with HCl, and evaporated that solution to dryness, it leaves at first a gelatinous mass, and after complete dryness, an insoluble gritty powder, it is *silica*.

For use, 1 part of the dry carbonate is dissolved in 4 or 5 parts of water, and kept in a bottle, closed with a gutta-percha stopper.

The principal applications of carbonates of potash or of soda *in the wet way*, are for the separation of the earths, and of a great many metallic oxydes which these reagents precipitate from their solutions, chiefly in the state of carbonates, though sometimes the precipitate is a hydrate of the oxyde, as is the case for example with alumina, and peroxyde of iron. In the reactions mentioned below, the precipitate, unless otherwise indicated, is a carbonate of the base.

The aqueous solution of carbonate of potash, or of soda, is used also for decomposing many organic salts, which, being boiled with it, are thus converted into a carbonate of the base of the salt, whilst the organic acid combines with the potash

or soda of the alkaline carbonate. It is used also for neutralising the free acid contained in certain liquors, and for decomposing certain salts, which being insoluble in water and in acids, are, however, decomposed when boiled with a solution of carbonate of potash or of soda. For example, BaO,SO_3 which is thereby converted into BaO,CO_2, &c.

In the dry way it is employed to decompose and render soluble those compounds which resist the action of water and of acids, such as for example certain silicates and earthy sulphates. In that case, however, it is better to use a mixture of both carbonates, in the proportion of about 10 parts of NaO,CO_2 and 13 parts of dry KO,CO_2 well mixed together, because this mixture fuses at a much lower temperature than either of them separately; the mixture should be kept in a well-stoppered bottle. Or the compound may be mixed thoroughly with carbonate of soda, the operator then presses the mass in the middle with the finger, and a piece of hydrate of potash being placed in the depression thus made, the whole is fused in the usual way, or over my gas lamp furnace.

In fusing insoluble sulphates, such as BaO,SO_3 or StO,SO_3 with an alkaline carbonate, the carbonic and the sulphuric acids transpose, and on treating the fused mass with boiling water, the soluble alkaline sulphate and the insoluble earthy carbonate produced may be easily separated by filtering, the reaction is as follows:

$$BaO,SO_3 + NaO,CO_2 = BaO,CO_2 + NaO,SO_3.$$

the insoluble carbonate remains on the filter, and the soluble sulphate passes through; but if the operator, instead of dissolving the fused mass in water, unguardedly treats it with HCl or another acid, with a view, as he may think, *to expedite* its solution, as I have read it erroneously recommended in some chemical works, retransposition takes place, and the operation is to be begun *de novo;* in effect, by pouring HCl for example upon the fused mass which, as we said, consists then of an earthy carbonate insoluble in water, and of an alkaline

sulphate soluble in water, carbonic acid is evolved, and the following retransposition takes place:

$$BaO,CO_2 + NaO,SO_3 + HCl = BaO,SO_3 + NaCl + HO + CO_2.$$

(See also Table XXVIII., A., Observations *l*, *m*.)

Carbonate of soda has almost the same action as KO,CO_2, except with solutions of peroxyde of platinum, which are not precipitated at all by that reagent whilst carbonate of potash produces in them a *yellow* precipitate; it is easily prepared by exposing the bicarbonate as we said to incipient redness for some time. As it is not contaminated by alumina or silica like KO,CO_2 it is often substituted for the latter on that account, but it may, and often does, contain all the other impurities, and also NaS—NaO,SO_2 or even NaO,S_2O_2, those impurities are detected as in the case of the potash salt; as to NaS its presence is detected by the odour of HS which is evolved by supersaturating by an acid (sulphuric acid for example) a portion of the salt previously dissolved in water. If a sulphite or hyposulphite be present, this supersaturation with an acid will evolve SO_2, recognisable by its odour of burning sulphur, and this odour, in the case of a hyposulphite, will be accompanied by a milkiness or a deposit of sulphur. Solution of acetate of lead should produce a perfectly *white, and not a brown, nor a black precipitate* in the solution of those carbonates.

The observations made a few lines above in reference to the treatment of the fused mass with water, must be attended to.

Carbonate of soda is also used with the blowpipe and charcoal as a flux and as a reducing agent. The mineral to be submitted to treatment should first be reduced to fine powder and mixed with from 6 to 8 times its weight of NaO,CO_2, the mixture is then placed in a hole scooped out in the charcoal support, and exposed first to the outer flame of the blowpipe. If the metal of the compound be thus reducible, it either appears as a fused globule, or else it sinks into the char-

coal and might escape detection, wherefore the surface of the charcoal should be scraped and triturated with water in an agate mortar, the crushed mass should be washed by elutriation and decantation, so as to remove the charcoal. The metal is then found in the agate mortar in spangles, which become quite bright by trituration. (See Table I., B., Observation *h*, *i*.)

Carbonate of soda is also employed with the blowpipe and a platinum hook, for the detection of chromium and of manganese, the first yielding a yellow, the second a turquoise blue bead. And also for the detection of mercury, See Col. 13, 16, 17, 21, and Table I., A., Observation *u*.

REACTIONS.

Lithia . . .	*Nothing* . .	in dilute solutions;
	White . .	scanty precipitate in concentrated solutions, and only after a long time.
Baryta . . .	*White* . .	precipitate; soluble, with effervescence, in acids.
Strontia . .	*White* . .	precipitate; soluble, with effervescence, in acids.
Lime . . .	*White* . .	precipitate; soluble, with effervescence, in acids.
Magnesia . .	*White* . .	precipitate; soluble, with effervescence, in acids; soluble also in NH_4Cl, but it reappears by boiling. (See Table VI., Observation *o*.)
Alumina . .	*White* . .	precipitate (Al_2O_3, HO); insoluble in an excess of the reagent.
Glucina . .	*White* . .	bulky, flocculent precipitate; soluble in a great excess of the reagent.
Thorina . .	*White* . .	precipitate; soluble in an excess of a concentrated solution of the reagent, but difficultly soluble if the excess of the reagent is a diluted solution thereof.
Yttria . . .	*White* . .	bulky precipitate; soluble in a great excess of the reagent.

Protoxyde of Cerium	*White* . .	bulky precipitate; sparingly soluble in an excess of the reagent.
Zirconia . .	*White* . .	bulky precipitate; soluble in a great excess of the reagent.
Protoxyde of Manganese .	*White* . .	precipitate; sparingly soluble in NH_4Cl.
Sesquioxyde of Manganese . .	*Brown* . .	voluminous precipitate (Mn_2O_3,HO).
Oxyde of Zinc	*White* . .	precipitate ($3ZnO,HO) + 2ZnO,CO_2$); insoluble in an excess of the reagent; soluble in KO and in NH_3.
		If much ammoniacal salt is present *no precipitate*, except by boiling.
Protoxyde of Cobalt	*Red* . . .	gelatinous precipitate, becoming *blue* by boiling.
Protoxyde of Nickel	*Apple-green*	precipitate; paler than that with KO.
Protoxyde of Iron .	*White* . .	precipitate (FeO,HO), turning
	Green . .	then
	Reddish-brown	at the surface.
		The *white* precipitate is soluble in NH_4Cl; but is reprecipitated by exposure.
Peroxyde of Iron .	*Reddish-brown* .	precipitate (Fe_2O_3,HO); of a lighter hue than that produced by KO.
Oxyde of Cadmium	*White* . .	precipitate; insoluble in an excess of the reagent.
Protoxyde of Lead	*White* . .	precipitate; insoluble in an excess of the reagent; soluble in KO.
Oxyde of Bismuth.	*White* . .	precipitate; insoluble in an excess of the reagent.
Protoxyde of Uranium . .	*Dingy-greenish*	precipitate; soluble in a large excess of the reagent.

Suboxyde of Copper	*Yellow* . .	precipitate.
Protoxyde of Copper	*Blue* . . .	precipitate; becoming black by boiling.
Oxyde of Silver .	*White* . .	precipitate; soluble in NH_3.
Suboxyde of Mercury . .	*Dingy-yellow*	precipitate; soluble in a great excess of the reagent; turning *black* by boiling.
Peroxyde of Mercury . .	*Reddish-brown* .	precipitate; insoluble in an excess of the reagent.
Protoxyde of Platinum . .	*Brownish* .	precipitate; and the liquor is *brownish* also.
Peroxyde of Platinum . .	*Yellow* . .	precipitate; especially with HCl ($PtCl_2$, KCl). CARBONATE OF SODA PRODUCES NOTHING.
Oxyde of Palladium	*Brown* . .	precipitate (PdO,HO); soluble in an excess of the reagent.
Peroxyde of Rhodium . .	*Nothing* . . *Yellowish* .	at first; after a time, precipitate (Rh_2O_3,HO).
Binoxyde of Iridium . .	*Reddish-brown* .	abundant precipitate; gradually redissolved, and the liquor becomes *colourless*.
Binoxyde of Osmium .	*Nothing* . . *Black* . .	at first; after a time, precipitate; and the superincumbent liquor is *bluish*; with NaO,CO_2, *the liquor is less bluish*.
Protoxyde of Tin .	*White* . .	precipitate (SnO, HO); insoluble in an excess of the reagent.
Peroxyde of Tin .	*White* . .	precipitate (SnO_2,HO), CO_2 is disengaged; the precipitate is soluble in excess, but reprecipitated by standing. The precipitate produced in perchloride of tin is insoluble in an excess of the reagent.

Oxyde of Antimony	*White* . .	voluminous precipitate; which settles after some time, and is soluble in an excess of the reagent. If *tartaric acid* is present, *no precipitate* is at first produced.
Protoxyde of Molybdenum	*Brownish-black* . .	precipitate (MoO,HO); slightly soluble in an excess of the reagent.
Binoxyde of Molybdenum	*Light-brown*	precipitate (MoO_2,HO); soluble in an excess of the reagent.
Binoxyde of Vanadium	*Greyish-white*	precipitate (VaO_2,HO); soluble in an excess of the reagent. The solution has a brown colour.
Oxyde of Chromium	*Light-green*	precipitate (Cr_2O_3,HO); almost *blue* by standing, and *violet* when viewed by artificial light: the superincumbent liquor is *green*.
Tellurous acid . .	*White* . .	abundant precipitate; soluble in an excess of the reagent.
Titanic acid	*White* . .	bulky precipitate (TiO_2,HO); insoluble in an excess of the reagent.
Antimonic acid . .	*White* .	precipitate; partly soluble in an excess of the reagent.
Antimonious acid . .	*White* . .	precipitate; partly soluble in an excess of the reagent.
Permanganic acid	*Brown* . .	precipitate (MnO_2).

CARBONATE OF SODA (NaO,CO_2). See CARBONATE OF POTASH.

CHLORATE OF POTASH.

(KO,ClO_5.)

Chlorate of potash is a somewhat considerable article of

commerce It forms anhydrous, flat and tabular crystals, which require about 20 parts of water for their solution. It is frequently contaminated by chloride of potassium, in which case a white precipitate is produced in its aqueous solution, by AgO,NO_5. It has scarcely any use as a test, but it is frequently employed for the purpose of destroying the organic substances which might interfere with the compounds to be analysed. (See Table I., D, and Observations.)

In concentrated solutions of *strontia* it produces a *white precipitate,* but in even moderately strong solutions, nothing; in the *neutral solutions* of protoxyde of iron it forms a *yellow* or a *brown* precipitate, according to the temperature of the solution.

Chloric acid is scarcely used at all as a reagent. It has, however, been employed sometimes as a test for potash, in the solution of which it forms a white crystalline precipitate, which is but sparingly soluble in cold water—but *perchloric acid* is even preferable for that purpose, since perchlorate of potash is still less soluble.

CHARCOAL.

The principal use of wood charcoal is for blowpipe experiments, for which purpose that made of soft wood should invariably be chosen. It should be free from knots and from pieces of bark, as it otherwise will crack and splinter under the impression of heat, and the mass operated upon may thus be projected and lost. Charcoal is used also for the detection of nitrates, chlorates, perchlorates and bromates, which deflagrate when heated with it; for the reduction of arsenious and arsenic acids, and of various metallic oxydes.

Animal charcoal is used for decolourising solutions, and also for separating various alkaloids, such as quinine, strichnine, &c.

CHLORIDE OF AMMONIUM.

(NH_4Cl, or NH_3,HCl.)

HYDROCHLORATE OF AMMONIA.—SAL AMMONIAC.

The sal ammoniac of commerce is generally pure enough, after recrystallisation, for the analytical chemist. Yet, if it contain iron (which is generally the case, and is immediately ascertained by its solution being rendered of a dingy green colour or precipitated black by NH_4S), it may be removed by adding a few drops of NH_4S to the solution in water, leaving the whole at rest for a few hours, filtering, adding as much HCl to the filtrate as is necessary to give it a feeble acid reaction, boiling until all odour of HS has completely disappeared, filtering, neutralising the acid filtrate with NH_3, and evaporating to the crystallising point.

Of course, sal ammoniac may at once be procured by neutralising exactly some pure HCl with NH_3.

The impurities of sal ammoniac are NH_4O,SO_3 which generally is not of any consequence—NaCl—NaO,SO_3—MgO, SO_3—NH_4Br—Metallic oxydes (chiefly iron) and Organic matter.

Sal ammoniac, or its solutions, should volatilise *completely* when ignited in a platinum capsule, and it should not be discoloured, precipitated, or rendered turbid by NH_4S, nor by BaCl.

The presence of bromine is detected by chlorine and ether in the manner described for the detection of bromine (see hydrobromic acid in the other Dictionary); that of organic matter by the carbonaceous residue left upon a platinum foil after igniting it thereon.

Sal ammoniac is chiefly used to prevent the precipitation of certain bases, such as *protoxyde of manganese, magnesia, oxyde*

of zinc, oxyde of nickel, oxyde of cobalt, or of certain salts, like *tartrate of lime,* by ammonia or the alkaline carbonates. It is also used for precipitating *alumina* or *oxyde of chromium* from its solution in caustic potash or soda; also for redissolving certain precipitates produced by other reagents in *magnesian solutions,* and separating therefrom or identifying the *ammonia-phosphate of magnesia* which is insoluble therein from other magnesian precipitates. It is used again as a test for *platinum,* in the solutions of which, provided they be concentrated, it forms a *yellow* precipitate which is characteristic, but chloride of potassium precipitates platinum more completely. Lastly, it serves to distinguish *sulphuric* from *selenic acid.* (See Table I., C, Observation *e.*) In solutions of vanadic acid NH_4Cl produces a white flocculent precipitate, insoluble in HCl.

For use as a test, dissolve 1 part of pure sal ammoniac in about 8 or 10 parts of water.

CHLORIDE OF BARIUM.

(BaCl.)

Chloride of barium may be obtained by dissolving carbonate of barytes in dilute HCl, filtering, evaporating to dryness, redissolving in water, filtering, and evaporating to crystallisation. It is, however, very easily procured in a perfectly pure state from all respectable dealers.

Nitrate and acetate of barytes may be prepared in the same manner by using nitric or acetic acid instead of HCl.

Chloride of barium is sometimes contaminated by StCl or StO,NO_5, in which case the salt on being treated by alcohol, and inflamed, will burn with a characteristic red flame. If it have a yellow colour, Fe_2Cl_3 may be present, in which case NH_4S will turn it black or brown. If it contains CaCl it will more or less deliquesce by exposure.

Chloride of barium must be completely soluble in water, and its solution must then be perfectly neutral to test papers. It must not be precipitated, discoloured, or rendered turbid by NH_3—HS, nor NH_4S, and an excess of dilute SO_3 poured into it should precipitate it so completely, that the filtrate therefrom should not leave a vestige of a residue by evaporation to dryness.

BaCl is chiefly used for the detection of SO_3, which it precipitates in the state of BaO,SO_3 perfectly insoluble in water, and in acids; and also as a means of separating the acids into groups, many of them producing, with baryta, salts which are insoluble in water, but soluble in acidified liquids.

For use as a test, dissolve 1 part of the crystals in about 10 or 12 parts of water. In the reactions mentioned below, the solutions, unless otherwise indicated, are assumed to be neutral.

REACTIONS.

Sulphuric acid		In acid and neutral solutions,
	White	precipitate; insoluble in water and in acids. (See Table VII., Observations *s*, *t*, Table XXI., Observation *g*.)
Sulphurous acid	*White*	precipitate, soluble in dilute HCl. (See Table VII., Observation *t*.)
Selenic acid	*White*	precipitate; insoluble in NO_5, but decomposed by HCl, especially by boiling, after which it is soluble.
Selenious acid	*White*	precipitate; soluble in acids.
Telluric acid	*White*	bulky precipitate; which, after some time, becomes granular and heavy
Iodic acid	*White*	precipitate; soluble in NO_5.
Phosphoric acid		In neutral solutions,
	White	precipitate; soluble in acids, and slightly in NH_4Cl.
Phosphorous acid	*White*	precipitate; soluble in HCl and in NH_4Cl.

Boracic acid	*White* . .	precipitate; soluble in a large quantity of water, in acids, and in NH_4Cl, and other salts of ammonia. (See Table I.—D., Observation *d*; Table VII., Observation *q—b'*, Table XXI., Observations *c*, *j*; Table XXII., A, Observation *h*.)
Molybdic acid	*White* . .	precipitate, soluble in a great excess of water, in NO_5, and in HCl.
Tungstic acid	*White* . .	precipitate; insoluble in water, and partly decomposed by acids.
Arsenic acid	*White* . .	precipitate; soluble in acids, and in NH_4Cl. (See Table XXI, Observation *c*.)
Arsenious acid . .	*Nothing*. .	in acid solutions, or with free AsO_3; but *after neutralising* with NH_3,
	White . .	precipitate, after a long time; soluble in acids. (See Table XXI., Observation *c*.)
Antimonic acid . .	*White* . .	precipitate.
Antimonious acid . .	*White* . .	precipitate, slightly soluble in water.
Osmic acid .	*Brown* . .	precipitate
Vanadic acid.	*Orange-yellow*	bulky precipitate, slightly soluble in water.
Chromic acid	*Lemon-yellow*	precipitate, soluble in NO_5.
Carbonic acid	*White* . .	precipitate, in neutral solutions.
Bromic acid .	*White* . .	precipitate, soluble in water.
Silicic acid .	*White* . .	precipitate, soluble in HCl.
Hydrofluoric	*White* . .	precipitate; soluble in HCl.
Iodic acid. .	*White* . .	crystalline precipitate, in concentrated solutions. (See Table XXII.—A., Observation *i*.)
Citric acid .	*White* . .	precipitate, in neutralised and concentrated solutions. (See Table XXII.—A., Observation *h*.)

Tartaric acid	*White* . .	precipitate; slightly soluble in an excess of the reagent; more soluble in dilute acids. (See Table XXII., A, Observation *h.*)
Oxalic acid .	*White* . .	precipitate; almost insoluble in water; soluble in acids. (See Table VII, Observation *b'*; Table XXII.—A, Observation *h.*)
Gallic acid .	*Black* . .	precipitate; sometimes immediately, but always after a few minutes, and the superincumbent liquor is of *bottle-green* colour.

CHLORIDE OF CALCIUM.

(CaCl.)

Chloride of calcium is easily prepared in a perfectly pure state by dissolving chalk in HCl, filtering the solution, and adding an excess of NH_3 to the filtrate, which is then kept for some time at a simmering heat and allowed to remain at rest for several hours, after which the whole is thrown on a filter. The filtrate is then heated to boiling, and solution of carbonate of ammonia is added as long as a precipitate is produced; the precipitate, which is carbonate of lime, is to be separated by filtering, washed thoroughly, until in fact the water used for washing has no longer any action on litmus paper, and the well-washed carbonate of lime is then gradually added to boiling HCl, diluted with about five times its bulk of water, until the acid is exactly neutralised; this point being reached the neutralised liquor is boiled again for about ten minutes, again filtered, and the filtrate, which is pure chloride of calcium, is kept for use.

Solution of CaCl should be perfectly neutral to test papers: it must not be precipitated or rendered turbid or discoloured by ammonia, for this would prove the presence of phosphate of lime, or of alumina, &c. Nor by NH_4S, which would indicate iron,

or some other metallic oxyde, nor should its solution evolve ammonia when mixed with potash or with caustic lime.

The uses of CaCl as a special reagent are rare, but as a general reagent it serves to separate several organic acids into groups, as has been seen in the Tables. Its behaviour with substances is very similar to that of BaCl. In the solid state, or in lumps it is used for absorbing moisture, and therefore for drying gases.

In the reactions mentioned below, the solutions, unless otherwise indicated, are assumed to be neutral.

REACTIONS.

Sulphuric acid	. . . . . .	In the free state, and in concentrated solutions, especially with addition of alcohol,
	White . .	precipitate; insoluble in HCl. In neutral and concentrated solutions,
	White . .	precipitate, without addition of alcohol.
Sulphurous acid .	. . . . .	In concentrated solutions, especially with the addition of alcohol,
	White . .	precipitate; soluble in HCl.
Iodic acid. .	*White* . .	crystalline precipitate.
Boracic acid .	*White* .	precipitate; soluble in a great quantity of water, and in NH_3 (See Table VII., Observation *v*)
Hydrofluoric acid + NH_3	*White* . .	gelatinous precipitate; so transparent that it can hardly be seen; but addition of NH_3 produces a
	White . .	bulky precipitate of CaFl, which may be identified as described in the tests for HFl and fluorides.
Molybdic acid	*White* . .	precipitate; soluble in a great quantity of water; in NO_5, and in HCl.
Phosphoric acid . .	*White* . .	precipitate; soluble in acids.
Phosphorous acid	. . . . . .	In neutral solutions,
	White .	precipitate, especially by boiling (See Table XXII.—A., Observation *j*.)

Tungstic acid	*White* . .	precipitate.
Arsenic acid	*White* . .	precipitate, in neutral solutions; soluble in acids, and in NH_4Cl.
Arsenious acid . .	*White* . .	precipitate, in neutral solutions; soluble in acids, and in NH_4Cl.
Tartaric acid	*White* . .	precipitate; soluble in an excess of potash; reprecipitated by boiling, and redissolved by cooling. (See Table VIII., Observation *c.*)
Racemic acid	*White* . .	precipitate, as with tartaric acid.
Carbonic acid	*White* . .	precipitate, in neutral solutions.
Oxalic acid	*White* . .	precipitate; insoluble in water; soluble in acids.
Tannic acid .	*White* . .	precipitate.
Meconic acid	*White* . .	precipitate.
Citric acid .	*Nothing* . .	in the cold;
	White . .	precipitate, by boiling. (See Table XXII.—A., Observation *j.*)
Malic acid .	*Nothing* . .	either in the cold or by boiling;
+ alcohol	*White* . .	precipitate, by adding alcohol. (See Table XXII.—A., Observation *k.*)
Gallic acid .	*Black* . .	precipitate.

CHLORIDE OF GOLD. (See TERCHLORIDE OF GOLD.)

CHLORIDE OF IRON. (See PERCHLORIDE OF IRON.)

CHLORIDE OF LIME. (See HYPOCHLORITE OF LIME.)

CHLORIDE OF LEAD.

($PbCl$.)

Chloride of lead is easily prepared by boiling protoxyde of lead (litharge) with HCl, evaporating to dryness, and dissolving the residue in water.

This solution is only used occasionally as a means of precipitating *silver* when lead is present at the same time.

CHLORIDE OF MAGNESIUM.

(MgCl.)

Chloride of Magnesium is prepared by boiling an excess of magnesia, or of carbonate of magnesia, with HCl diluted with about twice or thrice its bulk of water, the solution is then filtered and used, with the help of ammonia, for precipitating phosphoric acid from its solutions or from solutions of the phosphates. A solution of sulphate of magnesia mixed with ammonia answers the same purpose.

CHLORIDE OF MERCURY. (See PERCHLORIDE OF MERCURY.)

CHLORIDE OF NITRYLE. (See AQUA REGIA.)

CHLORIDE OF PALLADIUM.

(PdCl.)

Chloride of palladium is prepared by dissolving palladium in aqua regia with the help of a gentle heat. The aqua regia should consist of two parts of concentrated HCl, and one part of NO_5, a little of the latter acid being added from time to time, pending the dissolution, until all the palladium is taken up. The solution is then evaporated to dryness by means of a steam bath, in order to expel the excess of acid, and one part of the dry residue is then dissolved in 8 or 10 parts of pure water.

This reagent is employed for distinguishing alkaline Iodides from Bromides, a black precipitate being produced with the first, and nothing with the second.

CHLORIDE OF PLATINUM. (See Bichloride of Platinum.)

CHLORIDE OF POTASSIUM.

(KCl.)

Chloride of potassium is obtained by saturating a solution of carbonate, or of caustic potash with HCl.

The use of chloride of potassium, as a reagent, is very limited. In the concentrated solution of *platinum* it produces a characteristic *yellow* precipitate, and in those of *tartaric acid* a white crystalline precipitate.

Chloride of potassium precipitates platinum more completely than sal ammoniac.

CHLORIDE OF TIN. (See Protochloride of Tin.)

CHLORINE.

(Cl)

Introduce into a flask one part of peroxyde of manganese, and about four parts of ordinary muriatic acid, stir the mixture well with a glass rod or by agitation, and apply a gentle heat. It is absolutely necessary to stir the mixture before applying the heat, for otherwise the retort or flask will almost be sure to crack. The flask is provided with a disengagement tube plunging into a bottle containing water as cold as possible, in order to dissolve the gas evolved. The aqueous solution of chlorine so obtained must be kept in small well-stoppered bottles, and in the dark, otherwise it will undergo speedy decomposition, and a weak solution of HCl only will be left.

The solution should decolorise litmus paper and solution of sulphate of indigo; when it fails in doing so, it is a proof that the aqueous solution of chlorine has undergone decomposition.

As a reagent, chlorine water and chlorine gas are used for the detection of bromine, hydrobromic acid and bromides, and sometimes of iodine and iodides; also for the purpose of oxydising certain substances, such as, for example, converting sulphurous into sulphuric acid, proto-salts of iron into persalts of the same metal, and for the decomposition of organic substances, when they would interfere with the production of certain reactions; for which purpose an evolution of chlorine gas into the liquid containing them is preferable to the aqueous solution just spoken of.

Hypochlorite of lime may sometimes be used instead of chlorine water.

In the *dry way* chlorine gas is used to convert certain metallic sulphurets into their chlorides. This is done by placing those compounds in a tube through which a current of dry chlorine is passed, heat being applied externally, at the same time, to those compounds which are thus converted into chlorides, some of which are volatilised without decomposition, such are the chloride of tin, of antimony, and of arsenic (see my edition of H. Rose, Practical Treatise of Chemical Analysis, vol. ii. p. 298), while others are fixed, such are the chlorides of alkalies.

CHLORONITRIC ACID. (See Aqua Regia.)

NEUTRAL CHROMATE OF POTASH.

(KO,CrO_3.)

Neutral chromate of potash is easily prepared by dissolving a certain quantity of the bichromate of potash of commerce in

water, and adding gradually thereto as much solution of potash as is necessary to impart to it a very feeble alkaline reaction, evaporating to a suitable degree, and setting to crystallise.

This reagent, as found in commerce, may be contaminated by sulphate of potash, and by nitrate of potash.

If on adding BaO,NO_5 to the solution of the test, a precipitate is produced, which is not completely redissolved by NO_5, the insoluble portion is BaO,SO_3, and indicates, therefore, the presence of SO_3. If nitrate of potash be present, the salt deflagrates when thrown upon ignited charcoal.

Solution of bichromate of potash gives generally the same reactions as the neutral chromate; and as the bichromate is generally found in commerce in a very pure state, it is preferable to use it as a reagent instead of the neutral chromate, or else to prepare the latter from it by the method indicated at the beginning of this article.

As tests, both chromates are principally employed for the identification of lead and of baryta.

For use, dissolve 1 part of KO,Cr_2O_3 in about 10 parts of water.

REACTIONS.

Baryta . .	*Yellow* . .	precipitate; entirely soluble in excess of NO_5.
Strontia .	*Nothing*. .	at first; after a short time,
	Yellow . .	precipitate; soluble in free NO_5. N.B. Solution of bichromate of potash produces no precipitate even after a long time in solutions of salts of strontia.
Protoxyde of Lead .	*Yellow* . .	precipitate; insoluble in dilute NO_5; soluble in pure KO.
Oxyde of Bismuth	*Yellow* . .	precipitate; soluble in dilute NO_5.
Protoxyde of Copper .	*Reddish-brown* .	precipitate; soluble in NH_3, and in dilute NO_5.

Oxyde of Silver	*Crimson-red* precipitate; becoming *Reddish-brown*; soluble in dilute NO_5, in NH_3, and in a large quantity of water.	
Suboxyde of Mercury	*Red, or orange-red*	precipitate.
Peroxyde of Mercury	*Reddish-yellow*	precipitate.
Oxyde of Chromium	*Yellowish-brown*	precipitate; and if NH_3 be added, the superincumbent liquor is *yellowish-brown* also.
Hydrobromic acid, Bromides	. . . . . .	with SO_3HO, and heat applied, *hyacinth red fumes*, or *liquid*, which is Br.

CHROMATE OF STRONTIA.

(StO, CrO_3.)

Chromate of strontia is prepared by precipitating a solution of a soluble salt of strontia (for example the nitrate of strontia) by chromate of potash. The well-washed precipitate, is left to digest in a flask with distilled water, and thus kept for use. It is only employed to distinguish baryta, with which its solution produces a yellow precipitate from strontia, with which its solution, of course, produces none.

COPPER.

(Cu.)

We have already described copper in the other Dictionary. It is used as a test only in the state of foil, sheet, or bar, or of turnings.

It serves to detect mercury in the salts of suboxyde and of peroxyde of that metal. (See Mercury, in the other Dictionary).

Also, in conjunction with oil of vitriol, to detect the presence

of nitric acid and of nitrates. (See Nitric Acid, in the other Dictionary.)

To detect the presence of arsenic, as contrived by Reinsch. (See Arsenious Acid, in the other Dictionary.)

To analyse ammoniacal salts and nitrogenised organic substances, the nitrogen resulting from the decomposition being measured.

Copper is also used as a test for oxygen. (See Oxygen, in the other Dictionary.)

CREAM OF TARTAR. (See Bitartrate of Potash.)

CORROSIVE SUBLIMATE. (See Perchloride of Mercury.)

CYANIDE OF MERCURY.

(HgCy)

Cyanide of mercury crystallises in prisms, and is easily soluble in water, especially in boiling water, and its solution is, or should be, neutral.

It is not much used as a reagent except to detect palladium, in the neutral solutions of which it forms a yellowish white gelatinous precipitate. (See Table XXIV., Observation *a*.) This precipitate becomes quite white by standing, and is soluble in HCl.

For use as a test, dissolve 1 part of HgCy in about 15 parts of water.

CYANIDE OF POTASSIUM.

(KCy.)

Cyanide of potassium is prepared as follows :

Take a certain quantity of the commercial ferrocyanide of potassium, reduce it to powder, dry it well by means

of a gentle heat and stirring, and mix 8 parts of this anhydrous and well pulverised mass with 3 parts of dry carbonate of potash, which should be free from sulphate; triturate carefully the two substances together, so that the mixture may be as homogeneous as possible, and fuse it at a bright red heat in a covered crucible, or better still, in an iron pot provided with a cover, and keep it in the fire until the mass appears in a state of tranquil fusion. Remove the pot or crucible from the fire, and carefully decant its contents into a warm porcelain dish, so as not to disturb the iron, which has separated, or fallen to the bottom of the crucible, and which, without proper attention, would be poured along with the fused mass. Allow the mass to cool, break it into fragments, and keep it for use in a well-stoppered bottle.

Cyanide of potassium prepared in this manner is not pure, but contains cyanate of potash, which, however, does not interfere with its use. It should be perfectly white, and free from all particles of charcoal or of iron, and by treating it with cold water it should dissolve entirely, and yield a perfectly clear solution. It must contain no silica, nor alkaline sulphuret, and consequently the precipitate which it produces in solutions of lead salts should be perfectly white, and after previously supersaturating it with HCl, and evaporating it to dryness, the dry mass should redissolve completely in water, and yield with it again a perfectly clear solution. It must not, however, be kept in solution, but should be dissolved when required in 5 or 6 parts of water.

Cyanide of potassium is an exceedingly poisonous substance, which has an odour of hydrocyanic acid, due to its being slowly decomposed by the moisture and carbonic acid of the air.

Cyanide of potassium in the *wet way* precipitates most metallic oxydes from their solutions, but its principal application in the wet way is for the separation of cobalt from nickel, and of copper from cadmium as has been seen in the Tables.

In the *dry way*, cyanide of potassium, mixed with its weight of carbonate of soda, is one of the most powerful agents of reduction. The compound to be submitted to its action before the blow-pipe should be first reduced into very fine powder, incorporated with the mixture of cyanide of potassium and of carbonate of soda, placed in a hole scooped out of a piece of charcoal, and heated thereon by means of the blow-pipe. The mass fuses most readily, sinks into the pores of the charcoal, leaving the reduced metallic bead in the state of a pure and clean globule. (See Table XXVIII., A., Observation *m.*)

If the oxydes of copper, of tin, of antimony, of zinc, be projected into cyanide of potassium in a state of fusion, they are immediately reduced, even at a dark red beat.

ETHER.

The use of ether as a test in inorganic analysis is almost limited to the separation of free bromine from solutions which contain it, and the sulphuric ether of commerce is sufficiently pure for this purpose. (See the *modus operandi*, Table VII., Observations *m'*, *n'*.) It is employed, however, as a solvent of many organic substances.

FERROCYANIDE OF POTASSIUM.

($K_2Cy_3Fe,3HO$), (or C_6NH_3Fe+2K), (or K_2Cfy), or ($Cy_3Fe_2 3CyK$.)

Ferrocyanide of potassium, also called prussiate, or yellow prussiate of potash, is met with in commerce in a state of sufficient purity for analytical purposes. The principal impurity is KO,SO_3, which is easily detected by dissolving a portion of it in a large quantity of water, and testing with

BaCl, which, in that case, will produce the usual white precipitate of BaO,SO_3, insoluble in water and in acids.

For use as a reagent 1 part of the salt is dissolved in 10 or 12 parts of water.

Applications.—Ferrocyanide of potassium is employed chiefly to detect the presence of copper and peroxyde of iron; but many other metals are precipitated from their solutions by this reagent. In using it, a certain degree of caution must be exercised, because the free acids alone, if somewhat concentrated, and especially with the help of heat, may partially decompose the salt, and produce a white precipitate, which becoming blue by the contact of the air, may either lead erroneously to the belief that iron is present in the compound under examination, (whilst the precipitate and colouration is in reality due to the decomposition of the reagent just alluded to) or by an admixture of blue so far obscure the real colour of the precipitate produced, as to render doubtful the presence of the particular substance tested for. In testing with this salt liquors which contain a free acid, the latter should therefore be previously neutralised, or even supersaturated with ammonia, and then acidified again with acetic acid.

REACTIONS.

Baryta . . .	*White* . .	precipitate, with a tinge of *yellow* in concentrated solutions only.
	Nothing. .	in dilute solutions.
Strontia . .	*Nothing*.	
Lime . . .	*White* . .	precipitate, in concentrated solutions; it augments by standing.
	Nothing. .	in very dilute solutions.
Magnesia . .	*White* . .	abundant precipitate, after a while.
Alumina . .	*White* . .	precipitate, with the help of heat. The liquor acquires a blue tinge from the decomposition of the reagent.

Thorina . .	*White* . .	heavy precipitate; soluble in acids.
Yttria . . .	*White* . .	precipitate.
Protoxyde of Cerium . .	*White* . .	precipitate, in neutral solutions.
Zirconia . .	*White* . .	precipitate, in solution of sulphate and chloride of zirconium.
Protoxyde of Manganese	*White, or reddish-white, or orange-white.* .	precipitate, soluble in free acids.
Sesquioxyde of Manganese . .	*Greyish-green*	precipitate.
Oxyde of Zinc	*White* .	gelatinous precipitate, insoluble in HCl.
Protoxyde of Cobalt	*Green* . .	precipitate, becoming grey, insoluble in HCl.
Protoxyde of Nickel	*White* . .	precipitate, tinged with green, insoluble in HCl.
Protoxyde of Iron .	*White* *Light-blue* . *Dark-blue* .	precipitate; becoming and then by exposure. The precipitate is insoluble in acids.
Peroxyde of Iron .	*Dark blue* .	precipitate; insoluble in acids.
Oxyde of Cadmium	*White* . .	precipitate, with a slight tinge of *yellow*, soluble in HCl.
Protoxyde of Lead	*White* . .	precipitate.
Oxyde of Bismuth	*White* . .	precipitate, insoluble in acids.
Protoxyde of Uranium. .	*Brownish-red*	precipitate.
Peroxyde of Uranium . .	*Brownish-red*	precipitate.

Suboxyde of Copper . .	*White* . .	precipitate; becoming reddish-brown by exposure.
Protoxyde of Copper .	*Reddish-brown* . *Crimson* .	precipitate; in dilute solutions, colour, and then precipitate. (See Table IV., Observation *r*)
Oxyde of Silver	*White* . .	precipitate; turns bluish by exposure
Suboxyde of Mercury .	*White* . *Bluish* .	gelatinous precipitate; turning by exposure.
Protoxyde of Mercury .	*White* . .	precipitate, which is rapidly decomposed by exposure into soluble percyanide of mercury, and insoluble protocyanide of mercury, which turns blue by exposure.
Protoxyde of Platinum .	*Nothing.*	
Peroxyde of Platinum .	*Yellow* . .	precipitate
Protoxyde of Palladium . .	*Nothing* .	at first, after a while, *thick and firm jelly, of an olive colour.*
Binoxyde of Iridium. .	*Decolourised*	solution.
Peroxyde of Gold .	*Emerald-green.* .	colour, in perchloride of gold ($AuCl_3$).
Protoxyde [illegible] .	*White* .	gelatinous precipitate If it have a *reddish tint,* copper is present.
Peroxyde of Tin	*Nothing.* . *White* . . *Yellowish* .	at first; after a while, turbidness; and by standing, *jelly;* insoluble in HCl.
Oxyde of Antimony .	*White* . .	precipitate; insoluble in HCl. If *tartaric acid* be present, *nothing.*

Protoxyde of Molybdenum	*Reddish-brown*	precipitate.
Binoxyde of Molybdenum	*Brown*	precipitate.
Binoxyde of Vanadium	*Greenish-yellow*	precipitate.
Titanic acid	*Dark dingy green*	precipitate, in acid solutions.
Molybdic acid	*Brownish-red*	precipitate, in acid solutions; soluble in NH_3.
Vanadic acid	*Green*	precipitate; insoluble in acids.

FERROCYANIDE OF POTASSIUM PAPER. (See Test Papers.)

FERRICYANIDE OF POTASSIUM.

($K_3Cy_6Fe_2$), (or Cy_3Fe_2,3CyK.)

Ferricyanide of potassium is prepared by passing a current of gaseous chlorine through a dilute solution of ferrocyanide of potassium in water (1 part of the salt in about 10 of water), until a few drops of the liquor previously diluted being tested by solution of Fe_2Cl_3, fails to produce a blue precipitate. (Take care not to pass too much chlorine, which would destroy the ferricyanide produced.) When this point is reached, the liquor must be evaporated nearly to the crystallising point, and as much of a solution of KO,CO_2 is added thereto as to give it a feeble alkaline reaction. The liquor is then filtered whilst hot, and the filtrate in cooling deposits fine rhombohedric crystals of a deep red colour, almost insoluble in alcohol.

For use as a reagent 1 part of the crystals are dissolved in about 30 parts of water.

Ferricyanide of potassium is chiefly used as a test for protosalts of iron in solutions which contain peroxyde of iron at the same time. It may also be employed as a test for other metallic oxydes, but as is the case with ferrocyanide of potassium, the operator will recollect that if a strong acid be present a precipitate of Prussian blue may be formed at the expense of the reagent itself, hydrocyanic acid being disengaged at the same time.

The presence of free alkalies on the other hand, interfere with this reagent as with ferrocyanide of potassium.

In the precipitates produced by ferricyanide of potassium in metallic solutions, the 3 equivalents of potassium of the reagent are replaced by 3 equivalents of the precipitated metal.

REACTIONS.

Protoxyde of Manganese . .	*Brown* . .	precipitate; insoluble in free acids.
Sesquioxyde of Manganese .	*Brown* . .	precipitate; insoluble in free acids.
Oxyde of Zinc . .	*Orange-yellow*	precipitate; soluble in free HCl.
Protoxyde of Cobalt	*Reddish-brown* .	precipitate; insoluble in HCl.
Protoxyde of Nickel	*Greenish-yellow* .	precipitate; insoluble in HCl.
Protoxyde of Iron .	*Dark blue* .	precipitate; insoluble in acids.
Peroxyde of Iron	*Nothing* . .	The liquor only becomes a little darker —But if the slightest trace of FeO be present, a *dark blue* precipitate is produced.
Oxyde of Cadmium	*Yellow* . .	precipitate; soluble in HCl.
Oxyde of Bismuth .	*Pale yellow*	precipitate; soluble in HCl.

Protoxyde of Uranium . .	*Brownish-red* / *Nothing* .	precipitate, after some time, but at first.
Suboxyde of Copper . .	*Reddish-brown*	precipitate.
Protoxyde of Copper .	*Greenish-yellow* .	precipitate, insoluble in HCl.
Oxyde of Silver .	*Brownish-red*	precipitate, resembling much that produced by NH_3 in solution of Fe_2O_3.
Suboxyde of Mercury	*Reddish-brown* . / *White.*	precipitate; which, after some time, becomes
Peroxyde of Mercury .	*Yellow* . . / *Nothing*	precipitate, but in solutions of corrosive sublimate,
Peroxyde of Platinum .	*Yellow* .	precipitate
Protoxyde of Tin .	*White* . .	precipitate; soluble in HCl.
Oxyde of Antimony	*Nothing* . .	or if a turbidness is produced it disappears by adding a few drops of acid
Protoxyde of Molybdenum .	*Reddish-brown*	precipitate
Binoxyde of Molybdenum . .	*Brown* . .	precipitate.
Binoxyde of Vanadium .	*Green* . .	gelatinous mass
Molybdic acid	*Brownish-red*	precipitate; soluble in NH_3

FERRICYANIDE OF POTASSIUM PAPER. (See Test Papers.)

FLUORIDE OF CALCIUM. (Fluor-spar.)

Fluor-spar is generally used as a flux, especially in the treatment of the ores of copper and of lead.

It is used also in blowpipe experiments to identify sulphate of barytes, sulphate of strontia, and sulphate of lime, these three substances being the only compounds which fuse with fluor-spar into a limpid bead which is perfectly *clear* and *colourless when hot*, but which *on cooling becomes milk-white.* The fluor-spar should be pulverised.

FLUX.

Any substance is a flux which, being comparatively easy to melt when exposed to heat, can by being mixed and heated with another substance of a more refractory nature, induce, or promote its fusion, decomposition, or reduction. Compounds which have proved insoluble in water and in acids are rendered soluble in one or the other of these menstrua after fusion with the appropriate flux. Both the substance to be fused, and the flux with which it is to be fused, should previously be reduced to as fine a powder as possible, and thoroughly mixed so as to form a homogeneous mass. Platinum crucibles are generally used for the purpose, except when the reduction of a metal is apprehended, and in other cases which have been enumerated p. vii. of the Atlas. The flux generally used in the laboratory is carbonate of potash or of soda, in the proportion of 4 or 5 parts of either carbonate to 1 part of the well-pulverised compound. It is better still to employ a mixture of both carbonates, as

already stated p. 229. Cyanide of potassium is also a powerful reducing and desulphurising agent either alone or mixed with carbonate of potash or of soda. (See Cyanide of Potassium.)

Hydrate, carbonate, or nitrate of baryta, are also used as fluxes; but as the first, namely, hydrate of baryta, fuses at a low red heat, without losing its water of crystallisation, it is preferable to the two other salts for this purpose.

Borax is frequently used as a flux, especially in the assays of gold and of silver, and mixed with charcoal for the assays of iron and of tin. (See Borax.)

The principal other fluxes are—The Black Flux, which consist of

Bitartrate of potash (cream of tartar),	2 parts.
Nitrate of potash	1 part.

Pulverise, mix thoroughly, and deflagrate the mixture by throwing it, by small portions at a time, into a red hot crucible. The result is, in fact, carbonate of potash, mixed with finely divided charcoal.

White Flux :—

Bitartrate of potash } equal parts.
Nitrate of potash }

Pulverise, mix, and deflagrate the mass as directed for black flux.

Cornish Reducing Flux:—

Borax	3	parts.
Nitrate of potash .	4	,,
Cream of tartar .	10	,,

Another Flux :—

Borax	1	,,
Nitrate of potash .	2	,,
Cream of tartar .	3	,,

The latter is the flux generally used for the reduction of galena ores, the operation is practised as follows :—

Weigh separately 800 grains of the ore previously reduced into very fine powder, and 300 grains of the above flux also in powder. First pour into a wrought iron crucible a large portion of the flux, then the ore, and over that the rest of the flux. Fuse the whole in a good air furnace until no bubbling is any longer observed, withdraw the crucible and smartly tap it upon a plate of iron so as to determine the agglomeration and sinking of all the reduced lead to the bottom of the crucible. Decant now carefully the melted scories into one of the recesses of a brass ingot mould, and as soon as the melted lead is seen at the bottom of the crucible uncovered by the scories, push them (the scories) away on each side of the crucible with a small stick kept for the purpose, in a moist state, and then pour the melted lead in another recess of the brass ingot mould. Return the scories into the crucible, remelt them with an addition of carbonate of potash, and when well fused pour the whole mass quickly and at once into the recess of the ingot mould. After cooling a small button of lead will be found under the scories, which button is of course added to the mass first obtained.

GEORGINA PAPER. (See Test Papers.)

FORMIATE OF SODA. (NEUTRAL.)

($NaO, \overline{Fo}O, 2HO$ or $NaO, C_2HO_3, 2HO$.)

Neutral formiate of soda is a salt which is very soluble in water, but which may be obtained in prismatic crystals with a rhombohedral base, insoluble in alcohol. It is a very powerful reducing reagent. In the wet way it is used to reduce the noble metals, in the dry way, that is to say at a red heat, it reduces most of the other reducible metals, such as copper, lead, tin, antimony, nickel, cobalt, &c.

GOLD.

(Au.)

Gold is a soft metal of a beautiful yellow colour. It is the most malleable and the most ductile of all metals. Its specific gravity is 19·5, and it melts a little above the fusing point of silver. Neither air nor water have the slightest action upon it at any temperature, and the acids singly do not attack it, but the mixture of nitric and of hydrochloric acid, known under the name of *aqua regia*, dissolves it readily.

REACTIONS.

Nitric acid		To a liquid containing NO_5, or a nitrate, add HCl and a piece of gold leaf, heat being applied,
	The Gold dissolves	To ascertain whether any gold has dissolved, filter, and to the filtrate add a few drops of SnCl; it will produce a *purple* precipitate.
Hydrochloric acid / Chlorine + NO_5 . . .		To a liquid containing HCl, or a chloride, add NO_5, and a piece of gold leaf; apply heat,
	The Gold dissolves .	To ascertain whether any gold has dissolved, test the liquor with SnCl; a *purple* precipitate is produced.
Mercury + Zinc . . .		Put a few drops of the solution of the salt of mercury upon a piece of gold, and touch with a piece of zinc or of iron;
	A silvery stain .	is produced.
Bromates / Chlorates / Chromates		Behave with *Gold leaf* as nitric acid, &c., that is to say, a piece of gold leaf is dissolved in liquids which contain these substances; and to which HCl has been added.

GREY LITMUS PAPER. (See Test Papers.)

HYDRATE OF BARYTA.

(BaO, HO.)

The crystals which are deposited in cooling from a solution in hot water of pure baryta (see Baryta Water), are hydrate of baryta. These crystals fuse at a gentle heat without losing their water, and they are used in the same manner as carbonate of soda for the analysis of silicates by fusion, the baryta forming with the silicic acid a basic silicate of baryta, so that in treating the fused mass with water, filtering, supersaturating the filtrate with HCl, evaporating to perfect dryness, drenching the dry mass with strong HCl, and treating it with boiling water, the pure silicic acid is obtained in the usual form of a white, gritty, insoluble powder, and the oxydes of the compound in the state of chlorides in solution. The fusion may be performed in a platinum or silver crucible, yet hydrate of baryta is seldom used.

HYDRATED OXYDE OF BISMUTH. (See Oxyde of Bismuth.)

HYDRIODATE OF POTASH. (See Iodide of Potassium.)

HYDROCHLORATE OF AMMONIA. (See Chloride of Ammonium.)

HYDROCHLORIC ACID.

(HCl.)

Hydrochloric acid is of all acids that which is most frequently

used by the analytical chemist; but though found in abundance in commerce, it is seldom in a state of sufficient purity for analytical purposes, even that which is sold at a high price for the use of the chemist under the name of acid. hydrochlor. pur. is often contaminated by various impurities which render it unfit for use as a reagent. The ordinary or yellow muriatic acid of commerce is always impure, and contains sulphuric acid, iron, sulphurous acid, arsenic, nitric acid, free chlorine, and sometimes sulphate of soda, chloride of lead, likewise the salts which existed in the water which has been employed for condensing the gas, if pure distilled water have not been used. The substances which contaminate the acid. hydrochlor. pur. are generally arsenic, sulphurous acid, sulphuric acid, and nitric acid, or free chlorine.

Sulphuric acid is detected by adding a few drops of BaCl to a portion of the acid, in which case a white precipitate, *insoluble in water and in acids,* will be produced. Take care, however, before testing the acid with BaCl, to dilute largely with water, for BaCl always produces a white precipitate when poured in concentrated HCl; this white precipitate is nothing else than chloride of barium, which is insoluble in the concentrated acid, but which immediately dissolves on adding water. If the acid contains only traces of SO_3, several hours may be required for the precipitate produced by BaCl to appear.

Iron is easily detected by supersaturating a portion of the hydrochloric acid under examination with NH_3, which will precipitate the iron in the shape of reddish-brown flakes (Fe_2O_3), either immediately, or if in trifling quantity, after a little time. If, however, there are only traces of iron, it is better to supersaturate as just said with NH_3, and then to add NH_4S, which will either produce a black precipitate, or a green tinge, though this precipitate or tinge is sometimes due, not to iron, but to organic matter.

Copper is detected by supersaturating the acid with NH_3, adding an excess of acetic acid, and then a few drops of K_2Cfy. If copper be present, a crimson precipitate, or a pink colour, will be produced. If copper be present in sufficient quantity, the supersaturation of the acid with NH_3 alone will produce a blue colour or tinge.

If in saturating with NH_3, as just said, the acid becomes milky, *lead* is probably present.

Sulphurous acid is detected by boiling a portion of the acid with NO_5, which converts the SO_2 into SO_3, and then adding a solution of BaCl, which will produce of course a white precipitate insoluble in water and in acids (BaO,SO_3).

Sulphurous acid is also detected by passing a current of HS through the acid, for in that case a milk-white precipitate of sulphur is produced. If the precipitate so produced is *yellow*, arsenious or selenic acid may be suspected, if *black*, lead or copper.

There are various other means of testing for SO_2, but the most delicate is that proposed by Messrs. Fordos and Gelis, by which the smallest traces of SO_2 may be detected. It consists in pouring the acid under examination into a flask containing some metallic zinc, and passing the hydrogen which is evolved, through a solution of subacetate of lead ($3PbO),\bar{A}$), or better still, through a solution of oxyde of lead in KO; a black precipitate of PbS will then be produced; or else heat a portion of the suspected acid, and add thereto a little SnCl, and a few drops of CuO,SO_3; if sulphurous acid is present, black flakes of sulphuret of copper will be produced.

Arsenic is detected by passing a stream of HS through the acid under examination, in which case a *yellow* precipitate will be produced, the best and readiest way, however, consists in taking about one fluid ounce or two of the acid, diluting it with six or eight times its bulk of water, and boiling the whole with a strip of pure and perfectly clean copper, in the course of half an hour, but generally in the course of a few minutes, a

grey, metallic film of metallic arsenic will completely cover the copper. This film may be identified as arsenic by introducing the copper foil into a small glass tube closed at one end, and which need not be more than ¼ of an inch in the bore, and submitting it therein for a few moments to the heat of a spirit-lamp, when it will be observed that a white ring will have condensed in the cold parts of the tube, which white ring is arsenious acid in beautiful octahedrons, when viewed through a powerful magnifying lens, and which may be dissolved by a few drops of water, and tested with ammonia-nitrate of silver, which will produce a characteristic *yellow* precipitate.

A sample of hydrochloric acid obtained from one of the most respectable firms of this city, which I had occasion to analyse recently, yielded by this test 0·06 of arsenic contained in three fluid ounces of the acid.

Free *nitric acid* may be detected by scraping a quill-pen and boiling the little shavings with the acid under examination; the quill shavings will then become yellow, especially at the edges.

Or else the acid may be mixed with one-fourth of its bulk of concentrated SO_3,HO, and the mixture having become cold, a crystal of protosulphate of iron is then dropped into the liquid, which will assume a brownish colour at the point where it touches the crystal.

Free *chlorine* is detected by boiling with a few drops of sulphate of indigo, which in that case will become decolorised. Nitric acid, however, produces the same effect.

Free *chlorine* and *sulphurous acid* cannot exist together in the acid; if, therefore, sulphurous acid have been detected before, and the sulphate of indigo is bleached, it must be due to the presence of free nitric acid. Otherwise, and if the absence of nitric acid have been proved, it must be due to chlorine.

Free *chlorine* may besides be detected with still greater

certainty, by diluting a portion of the acid with water, adding thereto a solution of iodide of potassium, and then a little cold mucilage of starch, which will then produce a blue colour if free chlorine is present.

To resume: pure dilute hydrochloric acid should not be affected by—

BaCl, in which case Sulphuric acid or a soluble sulphate is present.

NH_3, „ Iron, copper, lead are present.

NH_4S, „ Iron, copper, lead are present.

K_2Cfy after supersaturation with NH_3, and then with acetic acid, serves to detect copper.

HS serves to detect sulphurous acid, arsenic

Zn serves to detect sulphurous acid

Cu serves to detect arsenic.

FeO,SO_3+SO_3,HO serves to detect free nitric acid or nitrates.

Quill scrapings serve to detect free nitric acid.

Sulphate of indigo serves to detect free chlorine.

And lastly, if on evaporating a portion of the acid to dryness, a residue be left, *it is an impurity*, pure hydrochloric acid being completely volatile. If that residue be carbonaceous, organic matter is present.

When hydrochloric acid contains *sulphurous acid*, the latter acid may be removed by adding a small quantity of peroxyde of manganese in very fine powder, and agitating the whole; a certain quantity of free chlorine is thus evolved, which transforms the sulphurous into sulphuric acid, thus:—

$$SO_2+HO+Cl= \\ SO_3+HCl.$$

By careful distillation the hydrochloric acid passes over, leaving the sulphuric acid in the retort. If the small quantity of *chlorine* contained in the acid, should prove objectionable, it may be removed by heating the acid

moderately, or by putting aside the first portions which come over in distilling, as long as the solution of sulphate of indigo is decolorised.

Perchloride of iron, and other fixed impurities, may be easily eliminated by redistillation.

If *arsenic* be present, it may be removed by distilling the acid with a small quantity of sulphuret of barium; three or four thousandths of the latter substance are generally sufficient to convert the arsenic into sulphuret of arsenic, which remains in the retort, and this addition may be made immediately after that of the peroxyde of manganese, as above described. *Arsenic* may also be completely removed by digesting the hydrochloric acid upon copper filings, decanting and then distilling the acid.

Howbeit, pure hydrochloric acid can always be easily prepared by mixing 13½ parts of oil of vitriol with 4 parts of water, the oil of vitriol *should be free from nitric acid*, otherwise the hydrochloric acid obtained will be contaminated by free chlorine, which, however, forasmuch as it is contained in the first portions which distil over, may be got rid of by keeping these first portions separate from the rest. The mixture of oil of vitriol and water having cooled, pour it upon 8 parts of the best common salt, contained in a mattrass or flask, carefully agitating the whole, so that the mass may be completely wetted, and then expose the retort containing the mixture to a moderate heat in a sand-bath; the retort should be provided with a disengagement-tube plunging about an eighth of inch into a receiver containing about 12 parts of pure distilled water, and kept as cool as possible by affusion of cold water. The acid so produced should be diluted with water, so that it may have a specific gravity of 1·11 or 1·12, such an acid containing then from 22·5 to 24·5 per cent. of pure hydrochloric acid gas.

Hydrochloric acid is mostly used as a solvent of a great many metals and oxydes which are insoluble in water. It is

generally necessary to dilute the acid with water and to apply heat, since otherwise no action may take place, for even certain carbonates fail in being decomposed by HCl, unless water be added and heat applied. (See Table I.—C, Observation *h*.) Hydrochloric acid is also used in preference to other acids for acidifying solutions.

The following metals are not acted upon by HCl:—

Antimony.	Platinum.
Copper (out of the contact of the air).	Rhodium.
Gold.	Silver (superficially).
Iridium.	Tantalum.
Lead (slightly).	Tellurium.
Mercury (slightly out of the contact of the air).	Titanium.
Molybdenum.	Uranium.
Osmium	Vanadium.
Palladium.	Zirconium.

All other metals are attacked by that acid, the hydracid being decomposed, hydrogen disengaged, and a chloride formed.

Certain peroxydes, for example, those of manganese and of lead, and other substances mentioned in Table I.—E, col. 6 ; several peroxydes and sulphurets treated by that acid are dissolved, the first disengaging chlorine, and the second sulphuretted hydrogen. (See Table XXIX., Observations *a*, *b*.)

Sulphites and hyposulphites, treated by that acid, disengage sulphurous acid.

As a special reagent HCl is used to detect *oxyde of silver*, *suboxyde of mercury*, and *lead*, and for the detection of ammonia in the presence of which it produces white fumes of sal ammoniac; for the latter purpose, however, strong acetic acid is a more conclusive test.

REACTIONS.

Protoxyde of Lead .	*White* . .	precipitate (PbCl); provided the liquor is not too dilute, the precipitate is soluble in KO, and in a large quantity of water; but NH_3 causes the white precipitate to reappear. (See Table XXIII., Observations *d, e.*)
Oxyde of Silver .	*White* . .	precipitate; insoluble in dilute acids, immediately soluble in NH_3, and re-precipitated by acids. (See Table XIV., Observation *b*; Table XVII., Observation *f.*; Table XXIII., Observations *d, e.*)
Suboxyde of Mercury . .	*White* . .	precipitate; insoluble in acids, turns *black* by NH_3. (See Table XXIII., Observation *d*)
Oxyde of Antimony .	. . . . . *White* . .	If tartaric acid is present, a small quantity of HCl produces a precipitate immediately soluble in a small excess of HCl.
Hyposulphuric acid . .	*Nothing* .	in the cold. By boiling, *odour of* SO_2, and the solution contains SO_3
Hyposulphurous acid . .	*Milkiness or Yellow* .	precipitate (S) after a time, accompanied by an odour of SO_2, and if nitrate of silver be added, the precipitate becomes *black*. (See Table VII., Observation *c.*)
Selenic acid .	*Nothing* . .	in the cold, by boiling, *odour of chlorine.*
Telluric acid, Chloric acid, Iodic acid .	*Nothing* . .	in the cold, by boiling, odour of chlorine.
Molybdic acid	. . . . . *White* . .	If the solution be a *molybdate of alkali*, precipitate, soluble in excess.

Tungstic acid	. . . . . *White* . . *Yellow.*	If the solution be a *tungstate of alkali,* precipitate ; becoming
Vanadic acid	*Odour of chlorine* .	and the solution dissolves gold and platinum.
Chromic acid	*Odour of chlorine* .	and the liquor becomes *greenish*
Permanganic acid.	*Odour of chlorine.*	even in the cold.
Manganic acid .	*Odour of chlorine*	and the *green* solution becomes *red,* and afterwards *colourless.*
Boracic acid .	*White* . .	precipitate. (See Table XXIII, Observations *d, e.*)
Hydrobromic acid, Bromides	*No apparent reaction* ,	but if a little bromate is present, the liquor becomes *yellow* or brown, from a liberation of bromine.
Hydriodic acid, Iodides .	*Nothing.* .	*apparently;* but if a little iodate is present, separation of *iodine;* and if heat is applied, *purple fumes* of iodine
Nitric acid and Nitrates .	. . . . .	dissolve gold leaf.
Benzoic acid .	*White* . .	precipitate. (See Table XXIII., Observations *d, e.*)

HYDROGEN.

Pure hydrogen is a colourless, tasteless, inodorous, and permanent gas; when prepared from zinc or iron, it always has a disagreeable, garlicky smell, which is due to the presence of traces of arseniuretted hydrogen, and of carburet of hydrogen. The odour is removed by passing the gas through a solution of potash and of AgO,NO_5. It is the lightest substance known.

Hydrogen is always prepared in the laboratory, by decomposing water by means of metallic zinc and of dilute sulphuric acid.

Hydrogen gas is used in analytical chemistry, for reducing certain metallic oxydes, sulphurets, and chlorides; for the detection of arsenic and antimony (Marsh apparatus); and for the analysis of atmospheric air.

HYDROFLUOSILICIC ACID.

($HFl, SiFl_2$) or ($2 Si Fl_3$) $3 HFl$)

Hydrofluosilicic acid is generally prepared by the operator as follows:—Introduce into a flask a mixture of equal parts of sand and of pulverised fluor-spar, pour upon the mixture about 6 parts of oil of vitriol, and adapt a disengagement-tube in such a way that it may plunge down a tall cylinder of glass full of water into a little mercury placed at the bottom thereof, in order that the orifice of the disengagement-tube may not be stopped up by the hydrated silicic acid which will separate. The disengagement of the gas begins even in the cold, but it should be promoted by applying the heat of a sand-bath to the flask. The liquid, which soon becomes gelatinous, should be occasionally stirred with a glass rod. When the evolution of gas has ceased, the gelatinous mass is squeezed through a piece of muslin, and then filtered. The filtrate contains hydrofluosilicic acid, and is kept as a reagent. It should produce no precipitate with salts of strontia.

It is used as a test for distinguishing baryta, in the solutions of which it produces a white precipitate, from lime and from strontia, in the solutions of which it produces none; and also sometimes as a test for potash, in the concentrated solutions of which it produces a gelatinous mass, almost white, and, in more dilute solutions, almost invisible.

REACTIONS.

Potash . . .	. . . .	In excess,
	Gelatinous .	*translucid, almost invisible,* precipitate; which, when dried, forms a white powder.
		In concentrated solutions,
	White . .	precipitate, or
	Turbidness.	
Soda		If the solution be not too dilute, and have been previously saturated with an acid, a
	Gelatinous .	*translucent* mass is produced.
Lithia . . .	*White* . .	precipitate.
Ammonia . .	*White* . .	abundant precipitate (SiO_3).
Baryta		After a short time,
	White . .	crystalline precipitate; almost insoluble in acids. (See Table VI, Observation *e.*)
Strontia . .	*Nothing.*	
Lime . . .	*Nothing.*	

HYDROSULPHATE OF AMMONIA. (See Hydrosulphuret of Ammonia.)

HYDROSULPHURET OF AMMONIA.

(NH_4S.)

Hydrosulphuret of ammonia is easily prepared by passing a stream of sulphuretted hydrogen (see the apparatus used for producing that gas under the head Sulphuretted Hydrogen), through an aqueous solution of ammonia. The ammonia must be perfectly saturated with the gas, and this is known when, by testing a portion of the liquor with solution of sulphate of magnesia, a precipitate is no longer produced.

The ammonia employed, if very strong, should be diluted with three or four times its bulk of water. It should be kept in well-stoppered phials free from lead, otherwise a black precipitate of sulphuret of lead will in course of time appear, which, however, may be separated by filtering.

Hydrosulphuret of ammonia should be perfectly transparent, when from long keeping it has turned of a deep yellow colour, and affords an extraordinary precipitate of sulphur when supersaturated with acids, without at the same time evolving a copious quantity of HS, it is no longer fit for use. It should completely volatilise by heat, and not be precipitated by solutions of salts of magnesia, even after standing, if it does, it is a proof that the ammonia is not perfectly saturated.

It is one of the most useful reagents, since it enables the operator to subdivide into groups those metallic oxydes which *are precipitated* from their acid solutions by HS, some of these precipitated sulphurets being soluble, whilst others are insoluble in hydrosulphuret of ammonia. Also to separate into groups those metallic oxydes which *are not precipitated* from their acid solutions by HS, but which are precipitated by NH_4S, whilst others are left unacted upon by either HS or by NH_4S. Most precipitates produced by NH_4S are sulphurets, except aluminium, zirconium, and chromium, which are precipitated in the state of hydrated oxydes. Wherefore, unless otherwise indicated in the reactions mentioned below, the precipitates produced may be considered as in the state of sulphurets. The substances or compounds which are precipitated from their solutions by NH_3, are likewise precipitated by NH_4S, and hence the arseniates, phosphates, borates of earths which are precipitated from their acid solutions by NH_3 are likewise precipitated by NH_4S, the only exceptions are the salts of magnesia, which though precipitated by caustic ammonia, yield no precipitate with pure hydrosulphuret of ammonia, except their solution is very concentrated.

REACTIONS

Baryta, Strontia, Lime, Magnesia	*Nothing*	but in the acid solutions of the *phosphates, borates,* and *arseniates of these earths,*
	White	precipitate.
Alumina		In neutral solutions,
	White	precipitate (Al_2O_3). (See Table V., Observation *f*, Table XVIII., Observations *d, i.*)
		In *phosphates, arseniates,* and *borates of alumina,* kept in solution by an acid, a white precipitate is produced. (See Table V., Observation *f*, Table XVIII., Observations *d, i.*)
Glucina	*White*	precipitate, in neutral solutions (GlO, HO), soluble in solutions of potash.
Thorina	*White*	precipitate (ThO,HO), in neutral solutions
Yttria	*White*	precipitate (YO,HO), in neutral solutions.
Protoxyde of Cerium	*White*	precipitate (CeO,HO). If a trace of iron or cobalt be present,
	Black	precipitate.
Zirconia	*White*	bulky precipitate (Zr_2O_3,HO); which, is
	Blackish	if there be a trace of iron
		If $\bar{T}$,2HO is present,
	Nothing	
Protoxyde of Manganese	*Flesh-red*	precipitate, *inclining to yellow,* in neutral solutions. It should be left to settle well, in order to judge of the colour (See Table V., Observation *f*; Table XVIII., Observation *d*.)
		The precipitate is
	Blackish	if a trace of iron be present.
Sesquioxyde of Manganese	*Flesh-red*	precipitate, in neutral solutions.

Oxyde of Zinc . .	*White* . .	precipitate; insoluble in an excess of the reagent, and in solution of the alkalies, and of their carbonates. (See Table V., Observations *f*, *v*; Table XVIII., Observations *d*, *i*.)
Protoxyde of Cobalt	*Black* . .	precipitate; insoluble in an excess of the reagent.
Protoxyde of Nickel	. . . . . . *Black* . .	In neutral solutions, precipitate; not quite insoluble in an excess of the reagent. The liquor remains *black* or *brown*.
Protoxyde of Iron .	. . . . . . *Black* . . *Reddish-brown*	In neutral solutions, precipitate; becoming by exposure; insoluble in an excess of the reagent
Peroxyde of Iron .	*Black* . . *Reddish-brown*	precipitate; insoluble in an excess of the reagent; becoming by exposure. If only a trace be present, the liquor assumes a *green colour*.
Oxyde of Cadmium	*Yellow, or Orange-yellow* .	precipitate, resembling orpiment, but not flocculent.
Protoxyde of Lead	*Black* . .	precipitate; insoluble in an excess of the reagent.
Oxyde of Bismuth.	*Black* . .	precipitate; if only a trace is present, a *dark brown colour or precipitate*.
Protoxyde of Uranium . .	. . . . . . *Black* . .	In perfectly neutral solutions, precipitate; insoluble in an excess of the reagent.
Peroxyde of Uranium . .	. . . . . . *Brown* . .	In neutral solutions, precipitate, insoluble in an excess of the reagent The superincumbent liquor first appears *black*.
Suboxyde of Copper . .	. . . . . . *Black* . .	In neutral or in alkaline solutions, precipitate; almost insoluble in an excess of the reagent.
Protoxyde of Copper . .	. . . . . . *Black* . .	In neutral solutions, precipitate, if only a small quantity is present,

	Brown . .	precipitate; slightly soluble in an excess of the reagent. (See Table XVI., Observation *b*; Table XVIII., Observation *b*, *c*.)
Oxyde of Silver. .	*Black* . .	precipitate; insoluble in an excess of the reagent, and in NH_3.
Suboxyde of Mercury . .	*Black* . .	precipitate; insoluble in an excess of the reagent, and in NH_3, insoluble in dilute acids (see Table XVII., Observation *a*); soluble in aqua regia, soluble in KO, but a *black residue* of metallic mercury is left.
Peroxyde of Mercury . .	*Black* . .	precipitate, insoluble in an excess of the reagent; insoluble in acids (see Table XVII, Observation *a*), soluble in aqua regia.
Protoxyde of Platinum .	*Black, or brownish-black* . .	precipitate; after saturating the liquor with KO, it has a *dark-brown colour*.
Peroxyde of Platinum . .	*Brownish-black* . .	precipitate; soluble in a pretty large excess of the reagent; and the solution has a *reddish-brown colour.*
Protoxyde of Palladium . .	*Black* . .	precipitate; insoluble in an excess of the reagent.
Peroxyde of Rhodium . .	*Brown* .	precipitate, insoluble in an excess of the reagent.
Binoxyde of Iridium. .	*Brown* . .	precipitate, soluble in an excess of the reagent.
Binoxyde of Osmium .	*Brownish-yellow* .	precipitate, insoluble in an excess of the reagent.
Peroxyde of Gold .	. . . . . .	In neutral solutions,
	Dark . .	precipitate; soluble in an excess of the reagent.
Protoxyde of Tin .	*Chocolate-brown.* .	precipitate; soluble in a great excess of the reagent.

Peroxyde of Tin	*Yellow*	precipitate; soluble in an excess of the reagent, and in NH_3—KO—KO,CO_2
Oxyde of Antimony	*Red*	precipitate; soluble in an excess of the reagent.
Oxyde of Molybdenum	. . . . . . *Brownish-yellow*	In solutions neutralised with ammonia, precipitate, soluble in an excess of the reagent.
Binoxyde of Molybdenum	. . . . . . *Brownish-yellow.*	In solutions saturated with NH_3, precipitate; soluble in an excess of the reagent.
Binoxyde of Vanadium	*Brownish-black*	precipitate, soluble in an excess of the reagent. The solution is dark purple.
Oxyde of Chromium	*Greenish*	precipitate (Cr_2O_3,HO), probably mixed with a little sulphuret of chromium. (See Table XVIII, Observation *d.*)
Selenious acid	*Lemon-yellow*	precipitate, soluble in an excess of the reagent.
Tellurous acid	*Brown*	precipitate; soluble in an excess of the reagent.
Titanate of Alkalies	. . . . . . *White*	In acid solutions, precipitate, HS is disengaged.
Molybdic acid	*Nothing* *Golden-yellow*	at first, after a while, the liquor becomes of a colour, and dilute acids produce then a *brown precipitate.*
Tungstic acid	*Nothing*	at first; but after a while, on adding an acid, a *light-brown precipitate* appears.
Vanadic acid	*Brown*	colour, and if an acid be then added, *brown precipitate,* soluble in an excess of the reagent. The solution has a purple colour.
Permanganic acid	*Flesh-red*	precipitate, if an excess of the reagent be added.
Manganic acid	*Flesh-red*	precipitate, if an excess of the reagent be added.

Osmic acid .	*Black* . .	precipitate; insoluble in an excess of the reagent.
Hydrocyanic acid .	*Blood-red* .	colour. (See Table VII., Observation *y'*.)

HYDROSULPHURIC ACID (HS). (See SULPHURETTED HYDROGEN.)

HYPOCHLORITE OF LIME.

(CaO,ClO.)

Hypochlorite of lime, known also under the name of *chloride of lime,* is now largely manufactured as an article of commerce; but it may be easily prepared in the laboratory, by passing a stream of chlorine gas through a milk of lime, or through layers of caustic lime pulverised, taking care that there be always an excess of lime, for otherwise, a certain quantity of the hypochlorite would pass into the state of chlorate of lime (CaO,ClO_5). For nse, the mass should be diluted with water, and filtered. The filtrate is a solution of hypochlorite of lime.

This reagent is sometimes substituted for chlorine as a test. (See Chlorine).

INDIGO. (See SULPHATE OF INDIGO.)

INFUSION OF GALLS. (See also TINCTURE OF GALLS.)

Infusion of galls is readily prepared by digesting 1 part of best blue galls, previously reduced to coarse powder, in 4 parts of water; after a few hours the mass is strained and squeezed through a linen bag, about 2 parts of common salt are then added to the expressed liquid, and the whole is filtered. The salt is added to prevent its turning mouldy.

Infusion of galls is chiefly used for detecting small quantities of *peroxyde of iron* in neutral solutions, in which case a black colour and precipitate are produced.

It is also used as a test for gelatine, in the solutions of which it forms a curdy or flocculent precipitate of a drab colour.

It is also used as a test for quinine and other organic substances.

It may further be employed as a test for several other metallic oxydes, as follows:—

REACTIONS.

Protoxyde of Iron	*Nothing* .	in neutral solutions, and provided no peroxyde of iron is present, otherwise a
	purple black	precipitate, augmenting by exposure, is produced.
Peroxyde of Iron	. . . . . .	In neutral solutions
	purple-black	precipitate *and colour* (Ink) which is rendered *brownish-black* by ammonia.
Protoxyde of Lead	*Dingy-yellow*	precipitate, in neutral solutions.
Peroxyde of Uranium	*Dark brown*	precipitate, in neutral solutions.
Oxyde of Silver	*Nothing* .	at first.
	Black . .	deposite of metallic silver after a while.
Suboxyde of Mercury	*Light yellowish*	precipitate.
Peroxyde of Gold	*Black* . .	precipitate (Au); especially with the help of heat, it then becomes *brownish yellow*.
Protoxyde of Tin	*Light yellow*	abundant precipitate.
Peroxyde of Tin	*Nothing* .	At first.
	Thick jelly .	after a while.

Oxyde of Antimony	*White, or slightly yellowish*	precipitate.
Peroxyde of Vanadium.	*Blue* . . .	colour, very intense, almost *black*, by standing a *voluminous black* precipitate is produced; the superincumbent liquor is *bluish*.
Tellurous acid . .	*Yellowish* .	precipitate.
Tantalic acid . .	*Orange-yellow*	precipitate.
Titanic acid .	*Reddish-orange* .	precipitate.
Vanadic acid.	*Bluish-black*	precipitate, after some time, in neutral solutions.
Zirconia . .	*Yellow* . .	precipitate.

IODIC ACID.

(IO_5.)

Introduce into a retort 5 parts of fuming nitric acid and 4 parts of iodine, and heat the mixture over a sand bath, taking care to wash down with the liquid of the retort any iodine which may sublime. The heat must be applied until no further action is observed, and the digestion must accordingly be continued for several hours; decant the liquid portion and evaporate to dryness. The white residue left is iodic acid. As a reagent it is used only to distinguish morphia from the other alkaloids, since this is as yet the only alkaloid which, in the solid state or in solution, possesses the property of decomposing iodic acid. The modus operandi consists in first mixing a portion of the iodic acid with cold mucilage of starch, and then testing with it the liquor or the solid suspected to contain morphia, which, if this alkaloid be present,

will at once become of a blue colour, the intensity of which is proportionate to the quantity of morphia.

Iodic acid mixed with starch is also used as a test for sulphurous acid and for sulphocyanides, the mixture assuming at once a blue colour when exposed to the fumes of sulphurous acid, or of hydrosulphocyanic acid. (See Table XXVII.—A, Observation *d.*)

IODIDE OF POTASSIUM.

(KI)

Iodide of potassium, sometimes called *hydriodate of potash,* is now largely manufactured for the wants of photography, and is therefore met in a pure state in commerce, yet it is often contaminated by *iodate of potash, chloride of potassium,* or *of sodium,* or other metallic *chlorides, carbonate of potash,* or by free caustic *potash,* and by *bromide of potassium*; carbonate of alkali and chlorides are, however, the more frequent impurities.

The presence of *carbonates* is detected by saturating portions of the substance with an acid, which will at once produce an effervescence, due to an evolution of CO_2. Iodide of potassium which is contaminated by a carbonate is not completely soluble in alcohol.

The presence of KO,CO_2 may also be detected by mixing a portion of the iodide with a milk of lime and filtering. If a small quantity of iodine be then added to the filtrate it will dissolve therein *without* imparting a brown colour to it.

The presence of free *potash* is detected by introducing into the solution of the suspected salt a minute quantity of iodine, which in that case will dissolve without imparting any colour to the liquid.

The presence of *chlorides* is detected by mixing a portion of the iodide of potassium with a small quantity of bichromate

of potash, putting the pulverised mass in a glass test-tube, adding some concentrated sulphuric acid, closing the tube with a cork provided with a disengagement tube plunging in a solution of potash. If now, heat being applied, a blood-red gas or liquid passes over, imparting a yellow colour to the potash liquid, a chloride is present. If, on the contrary, the products of distillation leave the potash solution colourless, no chloride is present.

The presence of chlorides may also be recognised by precipitating the solution of the iodide of potassium under examination by nitrate of silver mixed with a tolerably great excess of ammonia; iodide of silver alone is precipitated. The liquor is then filtered, and if on supersaturating the filtrate with NO_5, a white curdy precipitate (AgCl) is produced, insoluble in acids, and immediately soluble in excess of NH_3, then a chloride is present. It should, however, be borne in mind, that iodide of silver is *not quite insoluble in* NH_3, and therefore a slight turbidness may be produced by supersaturating the filtrate with NO_5.

If bromide of potassium is present, it is detected as directed in Table XXVI., B, col. 13.

Iodide of potassium is easily prepared by dissolving iodine in a solution of potash, the reaction yielding a mixture of iodide and of iodate of that base, thus—

$$6I + 6KO = \\ KO, IO_5 + 5KI.$$

The solution should then be evaporated to dryness, and the dry mass being mixed with a little charcoal powder, is ignited; the iodate under this treatment is decomposed into oxygen and iodide of potassium. The ignited mass is then treated by water, and the solution is evaporated to the crystallising point.

Iodide of potassium may also be readily obtained by boiling an excess of iron with water and iodine; the solution is

filtered and decomposed by adding carbonate of potash so as to precipitate the iron, which falls down in the state of carbonate of iron, and which is separated by filtering; by concentrating the filtrate crystals of iodide of potassium are obtained.

Solution of iodide of potassium is used as a test on account of the very characteristic colour of the precipitates which it forms in the solutions of several metallic oxydes; this reagent, however, is not much to be trusted for the detection of small quantities, because the precipitates which it produces are all more or less soluble in an excess of it.

It is principally employed as a test for *lead* and for *mercury*.

For use, dissolve 1 part of the salt in 10 parts of water.

REACTIONS.

Protoxyde of Lead	*Yellow*	precipitate, in neutral solutions; soluble in a large excess of the reagent. (See Table II.—B., Observation *h*.)
Oxyde of Bismuth	*Brown*	precipitate; very soluble in an excess of the reagent.
Suboxyde of Copper	*White*	precipitate (Cu_2I).
Protoxyde of Copper	*White*	precipitate; soluble in an excess of the reagent. The colour of the precipitate cannot be well seen except after separating it by filtering, the superincumbent liquor being coloured by free iodine.
Oxyde of Silver	*White*	precipitate, *with a yellow* tinge; very sparingly soluble in NH_3, soluble in an excess of the reagent, insoluble in dilute NO_5.
Suboxyde of Mercury	*Greenish-yellow*	precipitate; rendered *black* by an excess of the reagent, and soluble in that excess.

Peroxyde of Mercury . .	*Yellow* . .	precipitate; becoming of a beautiful
	Vermilion red . .	colour, immediately soluble in an excess of the reagent and in HCl. The dry precipitate yields by heat a sublimate which turns *red* by trituration. (See Table I.—A., Observation *r*.)
Protoxyde of Platinum . .	*Black* . .	precipitate.
Peroxyde of Platinum . .	*Dark brownish-red* .	colour, and after a time a *black* precipitate is produced, and the superincumbent liquor is colourless.
Binoxyde of Iridium . .	*Decolorisation* . .	of the solution, but *no precipitate.*
Binoxyde of Osmium .	*Nothing* .	at first. By long standing *black* precipitate, and the liquid becomes *bluish.*
Peroxyde of Gold .	*Black* . .	colour.
	Yellowish .	green precipitate.
Protoxyde of Tin .	*White* . .	curdy precipitate, with a slight *tinge of yellow.*
Peroxyde of Tin .	*Nothing.*	
Oxyde of Chromium .	. . . . . .	In neutral solutions
	Green . .	scanty precipitate; soluble in HCl.

IODINE.

(I.)

Iodine is used only in the state of saturated aqueous solution for the purpose of detecting the presence of starch in vegetable substances, and in the state of alcoholic solution for the detection of the alkaloids, with which iodine forms insoluble compounds.

IRON.

(Fe.)

Metallic iron precipitates the solutions of gold, silver, copper, tellurium, antimony, &c., in the metallic state. Bars of that metal, however, are chiefly and almost exclusively used for the detection of copper; a bar of clean iron, for example, the blade of a knife, or a needle, plunged in a solution which contains copper, becomes speedily coated over with a red film of metallic copper.

Iron is used also sometimes with boracic acid for the detection of phosphoric acid (see Boracic Acid, under which title the process is described); this test, however, is not of much value, and is accordingly but seldom resorted to.

LEAD.

(Pb.)

Pure lead obtained from the reduction of litharge or other oxyde of lead, or from a salt of that metal, is used for the analysis of silver in the dry way. The mineral containing silver is first roasted, if necessary, it is then mixed with 8 or 10 times its weight of pure lead, and a little NaO,CO_2, and fused upon charcoal before the blowpipe, or if in larger masses in a brasque crucible.

If the substance be in the metallic state, and consists of an alloy of silver with other metals, the mass is at once fused on charcoal with the same amount of pure lead as just said. The fused mass of lead, containing the other metals, is then placed upon a little bone cupel, which may be readily made by moistening bone ash with a little water, and kneading it into a stiff mass, which is then forced with moderate strength into

a pretty deep hole previously scooped out in a piece of charcoal, and pressed therein with a small agate pestle, so that the surface of the bone ash may be slightly concave. The bone ash cupel should then be slowly dried, and when dry, the little mass of fused lead is placed on it, and exposed to the outer flame of the blowpipe, so as to oxydise the lead which partly volatilises and partly sinks into the little bone cupel, nothing being left *on* the cupel but a small button, a small spangle of silver, which sometimes requires the help of a magnifying glass to be seen.

When larger masses are operated upon, a regular cupel is used, and the experiment is, of course, carried on in a muffle.

LEAD PAPER. (See Test Papers.)

LIME-WATER.

($CaO + Aq$)

Caustic lime or quicklime is only sparingly soluble in water, and more so in cold than in hot water; 100 parts of cold water dissolve only $\frac{1}{778}$ of lime.

Lime-water is easily prepared by slacking quicklime with water, shaking the hydrate so produced with a large quantity of distilled water in a bottle, allowing the undissolved portion to settle, and as the portion which has dissolved is generally contaminated by a certain quantity of potash, this first solution may be thrown away; a fresh quantity of water is then introduced, and after allowing the undissolved lime to settle, the clear superincumbent liquid is decanted and kept for use in well closed bottles.

Lime-water should impart to litmus paper a deep blue, and to turmeric paper a deep brown colour.

It is employed as a test for carbonic acid, for arsenious acid, for citric acid (by boiling), tartaric and racemic acids;

also for the detection of ammonia which is liberated from its salts by trituration with caustic lime in the solid state.

It is used also for the purpose of detecting phosphates.

REACTIONS.

Phosphoric acid . .	*White* . .	precipitate, in neutral solutions; soluble in acids.
Arsenic acid, Arsenious acid . .	*White* . .	precipitate, in neutral solutions, soluble in acids.
Carbonic acid	*White* . .	precipitate; soluble in acids, with effervescence. (See Table XXVII.—A., Observation *c.*)
Tartaric acid	*White* . .	In neutral solutions, precipitate; soluble in excess of $\bar{T}$,2HO, and in NH_4Cl.
Citric acid .	*Nothing* . .	in neutral solutions, in the cold; by boiling, *white* precipitate.

LIQUOR SILICUM. (See Silicate of Potash.)

LITHARGE. (See Protoxyde of Lead.)

LITMUS PAPER. (See Test Papers.)

MURIATE OF AMMONIA. (See Chloride of Ammonium.)

MURIATIC ACID. (See Hydrochloric Acid.)

MICROCOSMIC SALT. (See Phosphate of Soda and Ammonia.)

MORPHINE (CRYSTALLISED).

($C_{35}H_{20}NO_6 + 2HO$.)

Morphine or morphia is one of the alkaloids obtained from opium, and is one of the most active principles thereof. As a reagent it is only used for detecting or identifying nitric acid. Mr. O'Shaughnessy, who first proposed this test, gives the following directions:—Heat the supposed nitrate in a test tube with a drop of sulphuric acid and add a crystal of morphine, the liquor will then become *red* or *yellowish-red* if nitric acid is present. If a dilute solution be treated in this manner, it is necessary to apply heat, and the liquid, as it boils, acquires a yellow colour. Of course, the sulphuric acid employed must be free from nitric acid, which may be ascertained by testing it with morphia (or with brucine, see Brucine) in another tube.

MANGANATE OF POTASH.

(KO,MnO_3.)

The solution of green manganate of potash can be immediately prepared by fusing over a spirit-lamp in a glass test-tube a small quantity of peroxyde of manganese, nitrate of potash, and a small fragment of caustic potash; the whole on fusing forms a green mass, which, dissolved in water after it has cooled, forms the solution of green manganate. Dilute SO_3 may then be poured into this solution until it turns red. Such a solution is instantly decolorised by pouring it in a liquor which contains a sulphite or a hyposulphite, and is therefore used as a test for those acids (see Table VII., cols. 3, 4, Observation *g*); but HCl must not be present, since the latter acid decolorises also this manganate even in the cold.

NITRATE OF AMMONIA.

(NH_4O,NO_5.)

Nitrate of ammonia is prepared by saturating nitric acid with NH_4O,CO_2, and it is used for effecting the rapid combustion of organic substances in preference to nitrate of potash, because it is entirely volatile.

MOLYBDATE OF AMMONIA.

(NH_4O,MoO_3)

Molybdate of ammonia is prepared as follows:—Take the natural sulphuret of molybdenum, roast it until sulphurous acid fumes are no longer evolved, and digest the roasted mass (which is molybdic acid) with caustic ammonia, filter, and mix the filtrate with an excess of HCl, so as to redissolve the precipitate at first produced; the solution must be colourless.

Molybdate of ammonia is used only as a test for phosphoric acid. (See Phosphoric Acid in the other Dictionary.)

NITRATE OF BARYTA.

(BaO,NO_5.)

Nitrate of barytes is obtained by precipitating a solution of chloride of barium with an excess of carbonate of ammonia. The precipitated carbonate of barytes so obtained should be thoroughly washed, and then added to boiling dilute NO_5 to supersaturation. The hot liquid is then filtered, and the filtrate yields, by evaporation, perfectly pure crystals of nitrate of barytes.

Nitrate of barytes may also be prepared by treating the

native carbonate of barytes with dilute NO_5, evaporating to dryness, and purifying the salt by recrystallisation; but this process is longer.

Nitrate of barytes should be completely soluble in water, and its solution should be perfectly neutral to test-papers. It must not be discoloured or rendered turbid by NH_3,—HS—NH_4S, nor by AgO,NO_5, and an excess of SO_3 should precipitate its solution so completely that the filtrate, on being evaporated to dryness, should not leave the slightest residue.

Nitrate of barytes is used for the same purposes as chloride of barium, and is employed instead of the latter in those cases in which the introduction of a chloride to the liquid under examination would be either objectionable or inadmissible; see, therefore, Chloride of Barium.

For use, dissolve 1 part of the crystals in 10 of water.

NITRATE OF COBALT.

(CoO,NO_5.)

The solution of protonitrate of cobalt is prepared by dissolving pure oxyde of cobalt in nitric acid. It is, however, difficult to obtain any of the oxydes of cobalt in a state of perfect purity, they always retain traces of arsenic, of iron, and of nickel. The presence of arsenic is not of any consequence, and does not interfere generally with its use as a reagent, but it must contain no iron nor fixed alkali. Fresenius gives the following method of preparation, by which oxyde of cobalt sufficiently pure for the purpose may be obtained:—"An intimate mixture of 2 parts of very finely levigated cobalt ore, 4 parts of nitrate of potash, 1 part of effloresced carbonate of soda, and 1 part of dry carbonate of potash, are gradually projected into a red-hot crucible; the latter is afterwards left exposed to the strongest possible heat

until the mass is in an advanced state of fusion. The fused mass is subsequently allowed to cool, and afterwards levigated, and the powder boiled with water; the impure peroxyde of cobalt thus produced is thoroughly washed, subsequently digested and heated with hydrochloric acid until dissolved. The solution, which is of a dark-green colour, and gelatinous, owing to the separation of silicic acid, is evaporated to dryness, the residue is boiled with water and hydrochloric acid, filtered, and carbonate of ammonia added to the filtrate (which is kept boiling all the while) until all acid reaction ceases. The filtrate is then precipitated by means of carbonate of potash, the precipitate produced is *thoroughly* washed and subsequently dissolved in nitric acid. The solution is evaporated to dryness, at a gentle heat, and 1 part of the residue dissolved in 10 parts of water for use." The solution of protonitrate of cobalt produced in this manner is not quite free from contamination with nickel, but this contamination does not impair its fitness as a reagent.

The solution of nitrate of cobalt is used as a test for *magnesia*, for *alumina*, and for *oxyde of zinc*, in the following manner:—A small piece of the precipitate, or of the powder under examination, is heated red-hot upon charcoal before the blowpipe, and immediately moistened with a drop of the solution of nitrate of cobalt, and then it is strongly heated anew before the blowpipe. If, after the *complete cooling* of the mass, it has a pale-red or flesh colour, magnesia is present. If the colour of the mass is blue, *alumina* is probably present; we say probably, because a few other substances, for example, silica, give also a blue mass when so treated, this test being chiefly resorted to as a means of distinguishing magnesia from alumina. If the mass has a beautiful green colour, it is due to the presence of oxyde of zinc.

NITRATE OF LEAD.

(PbO,NO_5.)

Nitrate of lead is easily prepared by dissolving metallic lead or its carbonate in nitric acid, with the help of heat, filtering, evaporating the filtrate to the crystallising point, and allowing the liquor to crystallise. The crystals are anhydrous, of a milk-white colour, and opaque; they are in the form of octahedrons.

For use, dissolve 1 part of the crystals in 12 or 15 parts of water.

REACTIONS.

Hydrosulphuret of Ammonia	*Black* . .	precipitate.
Hydrosulphuric acid .	*Black* . .	precipitate.
Carbonates .	*White* . .	precipitate.
Chromates .	*Yellow* . .	precipitate; soluble in KO.
Bromates . .	*White* . .	precipitate; soluble in a large quantity of water.
Phosphates of Alkali	*White* . .	precipitate, soluble in NO_5.
Borates . .	*White* . .	precipitate, somewhat soluble in water
Molybdate of Alkali	*White* . .	precipitate.
Tungstic acid	*White* . .	precipitate.
Sulphuric acid and Sulphates	*White* . .	precipitate; almost insoluble.
Arseniate of Alkali	*White* . .	precipitate.

Osmic acid .	*Nothing* .	. at first; with addition of NH_3, *deep brown* precipitate.
Sulphurous acid—Sulphites	*White* .	. precipitate; soluble in cold NO_5; by boiling, nitrous acid fumes are produced, and sulphate of lead precipitated.
Iodic acid . .	*White* .	. precipitate, only after some time, and in concentrated solutions.
Hydrobromic—Bromides	*White* .	. precipitate; insoluble in water.
Hydrochloric—Chlorides	*White* .	. precipitate; soluble in a large excess of HO.
Hydrofluoric acid—Fluorides	*White* .	. precipitate.
Hydriodic acid—Iodides .	*Orange-yellow*	precipitate, soluble in hot water, and in NO_5.
Hydrocyanic acid—Cyanides.	*White* .	. precipitate.
Malic acid .	*White* .	precipitate, which becomes soft by boiling.
Oxalic acid and Oxalates . .	*White*	. abundant precipitate; sparingly soluble in $\bar{O}$,HO, soluble in NO_5, insoluble in NH_3.
Tartaric acid	*White* .	. precipitate, insoluble in HO; soluble in NH_3.
Citric acid .	*White* .	. abundant precipitate; sparingly soluble in NH_3.
Succinic acid.	*White* .	. precipitate, insoluble in an excess of the reagent.
Benzoic acid .	*White* .	. precipitate, in benzoates of alkalies, with benzoic acid the precipitate takes some time to appear.

NITRATE OF SUBOXYDE OF MERCURY.

PROTONITRATE OF MERCURY—SUBNITRATE OF MERCURY.

(Hg_2O, NO_5.)

Nitrate of suboxyde of mercury is prepared by putting in a beaker equal weights of mercury and of nitric acid of specific gravity 1·23, the whole being left at rest and *in the cold* for about twenty-four hours. At the end of that time the fine crystals which have formed are dissolved in water, acidified with NO_5, the liquor is filtered, and the filtrate kept for use in a bottle, the bottom of which is covered with metallic mercury.

The solution of subnitrate of mercury should contain no pernitrate, which is detected by precipitating with HCl, or a solution of common salt, and testing the filtrate with KI. If a red precipitate is then produced, it is a sign that a persalt of mercury is present.

The solution of subnitrate of mercury is chiefly used to precipitate gold and platinum, and as a test for ammonia, for formic acid, and other easily oxydisable substances.

REACTIONS.

Protoxyde of Platinum . .	*Black* . .	precipitate.
Peroxyde of Platinum . .	*Reddish-yellow* .	precipitate. (See Table IV., Observation *w*.)
Protoxyde of Palladium . .	*Black* . .	precipitate, in solutions of protochloride of palladium.
	Nothing. .	in other solutions of palladium.
Binoxyde of Iridium . .	*Light brown*	precipitate.

Binoxyde of Osmium .	*Yellowish white* . .	precipitate.
Peroxyde of Gold . .	*Black* . .	precipitate.
Ammonia . .	*Black* . .	precipitate.
Tetrathionic acid (S_4O_5) .	*Yellow* . . *Black* . .	precipitate. precipitate, if the salt be in excess.
Hyposulphurous acid . .	*Black* . .	precipitate is produced immediately.
Borates . .	*Olive-brown*	precipitate; soluble in NO_5 and in NH_3.
Molybdates of Alkalies . .	*Yellowish* .	precipitate; insoluble in HO; soluble in NO_5.
Hydrobromic—Bromides	*White* . .	precipitate; with a tinge of yellow.
Hydrocyanic acid—Cyanides .	*Grey powder*	Immediate reduction of metal in the state of grey powder, and the solution contains percyanide of mercury.
Chromic acid .	*Brick-red* .	precipitate.
Hydriodic acid—Iodides	*Yellowish green* . .	precipitate.
Bromic acid .	*Yellow* . .	precipitate; soluble in NO_5.
Oxalic acid	*White* . .	precipitate.
Tartaric acid .	*White* . .	precipitate.
Citric acid .	*White* . .	precipitate.
Lactic acid .	*White* . .	precipitate, in moderately concentrated solutions.
Succinic acid	*White* .	precipitate.
Acetic acid .	*White or yellowish* .	precipitate, in the cold; by boiling becomes *grey*, because mercury is revived.

Formic acid .	*White* . .	precipitate; soon becoming *grey* from reduction of mercury.
Malic acid .	*White* . .	Granular precipitate.

NITRATE OF NICKEL.

(NiO, NO_5)

Nitrate of nickel is prepared by dissolving pure nickel in nitric acid. The solution should, of course, be perfectly free from cobalt, which is easily ascertained by evaporating a small portion of the solution to dryness, and fusing a little of the dry residue with borax on the hook of a platinum wire, which, of course, will give a characteristic blue bead if cobalt be present.

The solution of nitrate of nickel is seldom used as a test, and then only as a means of identifying potash; because, according to Harkort, potash fused with borax in which pure oxyde of nickel has been dissolved, gives a bead of a bluish colour, or if too much nickel have been used, the bead has a dark-red colour.

NITRATE OF POTASH.

(KO, NO_5.)

To prepare pure nitrate of potash take a given weight of commercial nitrate of potash, dissolve it in its own weight of boiling water, and, after boiling the mass for a few minutes, filter it whilst still hot in a glass beaker, plunged in a larger beaker or in a basin full of cold water, and stir the whole filtrate with a glass rod until quite cold. The small crystals of nitrate of potash thus obtained are then thrown upon a filter and washed thereon with *cold* water until the liquor which filters is no longer rendered turbid when tested with nitrate of silver. The object of the stirring is to obtain the

nitrate in as small crystals as possible, because in that state they are more easily washed than when more voluminous. And if the original nitrate is very impure, it is best to dissolve and crystallise it a second, or may be a third time. When purified, the small crystals should be carefully dried between folds of blotting-paper.

Nitrate of potash should not be precipitated or rendered turbid by either nitrate of silver, nitrate of baryta, or carbonate of potash, otherwise the first indicates the presence of a *chloride;* the second, of a *sulphate;* the third, of a *metallic oxyde,* or of *an earth.*

Nitrate of potash is used in chemical analysis to detect the presence of carbon, and therefore of organic matters in substances which contain it, because, when thrown in nitrate of potash while in a state of fusion, deflagration takes place.

This salt is used also as a powerful means of oxydising a great many metallic sulphurets and other substances. In this way sulphurets of tin and of antimony are converted into oxydes of these metals, and sulphuret of arsenic into arseniate of potash.

Hyposulphates, hyposulphites, sulphites, and sulphates of metallic oxydes properly so called, thrown into nitrate of potash in a state of fusion, evolve nitrous acid fumes.

The salts or compounds of chromium thrown into fused nitre produce, after cooling, a yellow mass (KO,CrO_3) soluble in water, which thus acquires a more or less deep yellow colour.

Nitrate of potash is also used to destroy the organic substances present in a liquor, and which would interfere with the progress of analysis, as described in Table I.—D, and Observations thereon.

NITRATE OF SILVER.

(AgO,NO_5.)

The preparation of nitrate of silver is rapid and easy. It consists in taking an ordinary coin (a shilling or half-a-crown,

for example, which is an alloy of silver and copper) and dissolving it in nitric acid with the help of heat; the solution is evaporated to dryness, and the dry residue is transferred to a porcelain crucible, and kept in a state of fusion therein, at a moderate heat, until the nitrate of copper will have been completely decomposed and converted into black oxyde of copper, that is to say, until all traces of a green colour have vanished, and until nitrous acid fumes are no longer evolved, and therefore until a small sample thereof, being dissolved in water, fails in exhibiting a blue colour when tested with ammonia, or after complete precipitation by HCl and filtering, a crimson precipitate when tested with ferrocyanide of potassium. The fused mass is then allowed to cool, after which it is dissolved in boiling water, filtered to separate the black oxyde of copper, the filtrate is next concentrated to the crystallising point and left to crystallise.

For use as a test, one part of the crystals are dissolved in 20 parts of water.

The black oxyde of copper on the filter retaining still a pretty large quantity of silver should be redissolved in NO_5, and the silver is precipitated from this solution in the state of chloride of silver by adding HCl to it.

Nitrate of silver is now found in commerce in a perfectly pure state; but it sometimes happens that it contains *nitrate of copper*, or that it is adulterated with *nitrate of potash* and nitrate of lead.

Pure nitrate of silver should stand the following tests:— Its aqueous solution must not be rendered blue by ammonia, nor after complete precipitation by HCl must the filtrate turn pink or crimson by ferrocyanide of potassium, nor leave any fixed residue whatever, by evaporation to dryness.

Nitrate of silver is used as a general and as a special reagent. As a general agent, it serves to separate a whole group of acids, namely, PhO_5—$\bar{O}$,HO—BO_3—AsO_5—AsO_3—CrO_3—SiO_3—HCl—HBr—IO_5—$HCyS_2$—BrO_5—HI—

$HCy2S_2$; and again these into three classes, namely; the first seven being soluble in NO_5, the next five, being insoluble in this acid, but soluble in NH_3, and the last two, being insoluble both in NO_5 and in NH_3; and as nitric, chloric, perchloric, and acetic acids are not precipitated by AgO,NO_5, this reagent, therefore, is used to separate them into another group.

As a special reagent it is employed principally for the detection of HCl or of chlorides, and of other acids which form with silver precipitates very sparingly soluble or altogether insoluble in water, such as bromic, iodic, phosphoric, boracic, arsenic, and arsenious acids, hydrobromic acid and bromides, hydriodic acid and iodides; also, to detect or confirm the presence of chromic acid, and of formic acid.

REACTIONS.

Protoxyde of Iron	*White* . .	precipitate, in neutral solutions (Ag). If the salt of iron be in excess, the precipitate becomes black.
Tetrathionic acid.	*White* . .	precipitate, at first, but which soon becomes
	Yellow . .	and finally turns
	Black.	
Hyposulphurous acid . .	*White* . .	precipitate (AgO,S_2O_2), at first; but which after a while becomes yellowish-brown, and finally turns black, especially by boiling. (See Table VII, Observation *h*)
Vanadic acid.	*Yellow* . .	precipitate; becoming white by exposure, soluble in NO_5 and in NH_3.
Silicic acid .	*Yellow*	precipitate.
Bromic acid .	*White* . .	precipitate, soluble in NH_3, almost insoluble in dilute NO_5.
Iodic acid. .	*White* . .	precipitate; soluble in NH_3; but not in dilute NO_5.
Phosphate of Alkali	*Yellow* . .	precipitate, in neutral solutions; soluble in NH_3 and in NO_5. (See Table VII., Observations *y, z, c'*.)

Pyrophosphate	*White* . .	precipitate; soluble in excess of NO_5 and in NH_3. (See Table VII., Observations *y*, *z*, *c'*.)
Phosphorous acid	*Blackish-brown* .	precipitate.
Boracic acid		In concentrated solutions
	White . .	precipitate; soluble in a large excess of water. In dilute solutions
	Brown . .	precipitate (AgO), insoluble in water. (See Table VII., Observation *c'*.)
Molybdic acid	*White* . .	precipitate, soluble in a great excess of water; soluble in NO_5 and in NH_3.
Tungstic acid	*White* . .	precipitate.
Arsenic acid		In neutral solutions
	Light-brown	precipitate, very soluble in acids and in NH_3. (See Table VII., Observation *n*. Table XVI., Observation *c*.)
Arsenious acid		In neutral solutions
	Yellow . .	precipitate; very soluble in acids and in NH_3. (See Table VII., Observation *n*. Table XVI., Observation *c*.)
Sulphurous acid . .	*White* . .	precipitate, soluble in a large excess of sulphite, and which turns black by boiling. (See Table VII., Observation *h*.)
Bromic acid .	*White* . .	precipitate; soluble in NH_3 and in NO_5; but with difficulty; deflagrates when heated on charcoal.
Oxalic acid .	*White* . .	precipitate, soluble in NO_5 and in NH_3. (See Table VII., Observation *c'*.)
Hydrocyanic acid	*White* . .	precipitate; insoluble in NO_5; but soluble in NH_3. (See Table VII., Observation *g'*, Table XXVII.—B, Observation *a*.)
Hydrobromic acid	*Yellowish-white* . .	precipitate, insoluble in NO_5; soluble, but with difficulty, in NH_3. (See Table VII., Observation *g'*; Table XXI., Observation *v*.)

Hydriodic acid, Iodides .	*Yellowish-white* . .	precipitate; insoluble in dilute NO_5. and almost insoluble in NH_3.
Hydrochloric acid, Chlorides . .	*White* . .	curdy precipitate; immediately soluble in NH_3, insoluble in acids; fuses without decomposition. (See Table VII., Observations *g′*, *l′*.)
Hydrosulphuric acid . .	*Black* . .	precipitate.
Hydroferrocyanic acid . .	*White* . .	precipitate, insoluble in NH_3; becomes *brown* by boiling. (See Table VII., Observation *i′*.)
Hydroferricyanic acid . .	*Reddish-brown* . .	precipitate; soluble in NH_3, in the cold, reprecipitated by boiling. (See Table VIII., Observation *i′*.)
Chromic acid.	*Reddish-brown* .	precipitate, soluble in NO_5 and in NH_3.
Citric acid .	*White* .	precipitate, in neutral solutions; soluble in excess of NH_3.
Malic acid . .	*White* . .	precipitate, in neutral solutions.
Succinic acid	*White* . .	precipitate; immediately soluble in NH_3.
Benzoic acid .	*White* . .	precipitate; immediately soluble in NH_3.
Formic acid .	*White* . .	precipitate; soon becoming blackish from reduction of the silver.
Acetic acid .	*White* . .	crystalline precipitate, in neutral solutions, almost insoluble in cold water, more soluble in hot water.
Tartaric acid		In neutral solutions
	White . .	precipitate; insoluble in HO; soluble in NH_3.

NITROCHLORIC ACID. (See CHLORO-NITRIC ACID.)

NITROMURIATIC ACID. (See CHLORO-NITRIC ACID.)

NITROPHENISIC ACID. (See PICRIC ACID.)

NITROPICRIC ACID. (See PICRIC ACID.)

NITRIC ACID.

(NO_5,HO.)

Ordinary or commercial nitric acid is seldom pure; the substances by which it is principally contaminated are *hydrochloric acid* or *chlorine*, the presence of which is easily detected by means of solution of nitrate of silver, which, in that case, will produce the well-known white curdy precipitate of chloride of silver, immediately soluble in the slightest excess of ammonia. *Sulphuric acid*, the presence of which is detected by *diluting* a portion of the nitric acid under examination *with water* and testing with BaCl or BaO,NO_5, which, in that case, will produce a white precipitate of BaO,SO_3, insoluble in water.

Before testing the acid as just said, it is important not to omit to dilute it with three or four times its bulk of distilled water, for otherwise a precipitate would be produced; such a precipitate is nothing else than nitrate of silver or nitrate of baryta, which might thus simulate the presence of chlorine or of sulphuric acid. These precipitates, however, immediately disappear by adding water.

When nitric acid contains *fixed substances*, they are easily detected by evaporating a small quantity of the acid in a capsule, or on a strip of platinum foil, in which case a residue will be left; whereas nitric acid evaporates completely, without leaving a trace of fixed matter.

Nitric acid, however, can be easily obtained in a pure state from the commercial acid, by pouring into it a solution of

nitrate of silver as long as a precipitate is produced, adding to the liquid decanted from the precipitate a little pure nitrate of potash, in order to decompose any sulphate, or arrest any sulphuric acid which may be present, and then distilling in a retort almost to dryness.

Nitric acid can also be obtained perfectly free from HCl by simply distilling it, and collecting the acid for use only after one-fourth of it has distilled off. This first portion contains all the HCl, and may be employed for making *aqua regia.*

Pure nitric acid is colourless, but it has often a yellowish or a ruddy colour, due to the presence of hyponitric acid; this latter acid, however, does not interfere with the use of this reagent. Hyponitric acid may, besides, be removed by boiling.

Nitric acid contains also sometimes a little iodine, from the iodide of potassium which is contained in the nitrate of soda, from which it may have been manufactured. The presence of iodine is detected by diluting a portion of the acid with water, pouring into it a little mucilage of starch, and then sulphurous acid, drop by drop, when the blue colour of iodide of starch will be produced. If the nitric acid contains any nitrous acid, the starch will turn blue, without any addition of SO_2.

The uses of nitric acid are completely enumerated in my edition of Rose's "Practical Treatise of Chemical Analysis," in the following terms:—

"Nitric acid is employed in certain cases for dissolving the oxydised substances which are insoluble in water, when the presence of hydrochloric acid must be avoided. Although most of the combinations which it forms with the bases are soluble, it is by no means preferable to hydrochloric acid as a solvent of oxydised substances; because they generally dissolve less readily in nitric than in hydrochloric acid, and the excess of acid which it is then necessary to employ is more difficult to eliminate by heat.

"When operating upon non-volatilisable substances, the most troublesome peculiarity attending the use of nitric acid is the expulsion of the nitrate of ammonia, which is formed in large quantity when nitric acid is used, instead of hydrochloric acid, for dissolving an oxyde, and ammonia employed for supersaturating the solution thus obtained. This expulsion of the ammoniacal salt is attended with difficulties, because, in that case, when nitrate of ammonia exists in sufficient quantity along with organic substances, an explosion may often take place.

"Nitric acid is sometimes employed instead of hydrochloric, and especially for acidifying neutral or alkaline solutions, but this is seldom necessary.

"Nitric acid, however, is especially employed for dissolving metals and metallic alloys, because it often happens that metals cannot be dissolved by any other acid. Nitric acid is used also for oxydising the metallic sulphurets, nitrous fumes being evolved, and the liquor then contains sulphuric acid, due to the oxydisation of the sulphur of the sulphuret. It serves also to convert substances in solution into substances of a higher degree of oxydisation; for example, protoxyde into peroxyde of iron; tin, or the protosalts of tin, into peroxyde of tin.

"For ordinary purposes, the pure nitric acid employed is diluted with water to the strength of ordinary aquafortis, namely, to a specific gravity of 1·10. The cases in which nitric acid containing nitrous acid is employed are rare; such an acid, however, is used to oxydise certain metallic sulphurets. Instead of nitric acid only, a mixture of 1 part of nitric acid and 2 parts of hydrochloric acid (aqua regia) is sometimes employed, in which case it is of course immaterial that the nitric acid contain hydrochloric acid, but it must be free from sulphuric acid.

"Nitric acid, especially with the help of heat, dissolves nearly all the metals under disengagement of nitric oxyde,

and sometimes also of nitrous acid gas." It has, however, no action upon *gold, platinum, chromium, tungsten, tantalum, titanium, cerium, osmium, rhodium,* and *iridium.* Poured upon *lead, bismuth, copper, mercury, silver, palladium,* nitric oxyde is evolved, which, in contact with the air, forms ruddy fumes, and the result is a metallic nitrate, thus :—

$$3Pb + 4NO_5HO =$$
$$3PbO,NO_5 + NO_2 + 4HO.$$

Poured upon *arsenic, molybdenum, vanadium, antimony,* nitric oxyde is also evolved, but the metal is transformed into a metallic acid.

Poured upon *tin,* the metal is easily oxydised, nitrogen, nitrous acid and nitric oxyde gas being disengaged; nitrate of ammonia is formed, and insoluble peroxyde of tin remains (metastannic acid).

Poured upon *potassium,* and the metals of the first and of the second sections, upon *zinc, cadmium, iron* of the third section (see the Atlas, p. 4), nitrogen, or nitrous or nitric oxyde gas is disengaged, and nitrate of ammonia and a metallic nitrate are formed. Nitrate of ammonia is produced only when a certain quantity of water has been decomposed.

It dissolves nearly all the oxydes, except *peroxyde of tin, oxyde of antimony, tellurous acid,* and a few others, and also the salts produced by the oxydes which are insoluble in water.

Peroxydes are partially converted by nitric acid into a basic oxyde, and an oxyde of a higher degree of oxydisation.

The simple non-metallic bodies, such as *sulphur, selenium,* &c., are oxydised more easily by fuming than by dilute nitric acid, their combinations with the metals are likewise dissolved by nitric acid, but generally the metal is dissolved much sooner than the substance with which it is combined.

Sulphuret of mercury is almost the only metallic sulphuret which is not decomposed by digestion with nitric acid, aqua regia, however, readily attacks it.

Chloride, bromide, iodide, and *cyanide of silver, bromate, iodate of silver,* and *sulphates* and *seleniates of baryta, of strontia, of lime,* and *of oxyde of lead,* are insoluble in nitric acid. (See Table II.—B, Observations *c* and *e*.)

Fuming nitric acid is used in preference to ordinary nitric acid in certain cases, for the purpose of oxydising certain metallic sulphurets, but a more dilute acid (specific gravity, 1·11 or 1·12) is more frequently employed.

REACTIONS.

Protoxyde of Iron .	. . . . *Dark brown*	With the help of gentle heat colour, which vanishes by an excess of NO_5, or of FeO, or by exposure. (See Table V., Observation *h*. See also Perchloride of Iron, and Protosulphate of Iron.)
Oxyde of Antimony	*White* . .	insoluble precipitate; soluble in an excess of $\overline{T},2HO$. If an excess of $\overline{T},2HO$ exists in the original liquor, of course nitric acid produces *nothing*.
Protoxyde of Tin.	*White* .	precipitate; insoluble in the excess of the reagent (See Table II.—B, Observation *e*)
Hyposulphates	*Nothing* .	in the cold; by boiling, *ruddy fumes* of nitrous acid, and the liquor contains SO_3.
Sulphites . .	*Ruddy fumes,* but no deposit of sulphur.	
Hyposulphites	*Ruddy fumes* and deposit of S.	
Tetrathionic acid .	*Yellow* . .	precipitate, (S).
Molybdic acid	*White* . .	precipitate, in moderately concentrated solutions, soluble in the excess of the reagent
Tungstates of Alkali	*White* . .	precipitate, becoming yellowish.

Vanadic acid .	*Blue colour*	if the NO_5 be very concentrated and fuming.
Permanganic acid .	*Brown* . .	precipitate of peroxyde of manganese, with the help of heat.
Sulphurets .	*Ruddy* . .	fumes of nitrous acid, especially with the help of heat, sulphur being ordinarily separated, which, eventually, by boiling agglomerates into yellow lumps. (See Table I.—E, Observations *n*, *o*, *p*, *r*, *s*. Table II.—B, Observation *i*.)
Hydrobromic acid, Bromides	. . . . . *Nothing* ; . *Hyacinth red*	In the cold almost but by heating the mixture fumes of bromine appear, and the liquid can attack gold like aqua regia.
Hydriodic acid, Iodides . .	*Purple fumes* of iodine, but *heat must be applied.*	

NITROCHLORIC ACID. (See AQUA REGIA.)

NITROMURIATIC ACID. (See AQUA REGIA.)

NITROPHENISIC ACID. (See PICRIC ACID.)

NITROPICRIC ACID. (See PICRIC ACID.)

OXALATE OF AMMONIA.

($NH_4O,\bar{O}$.)

For the purposes of qualitative analysis, oxalate of ammonia may always be readily prepared by adding a slight excess of ammonia to oxalic acid. Oxalate of ammonia is preferable to oxalic acid, because the aqueous solution of the latter suffers decomposition by keeping, which is not the case with that of oxalate of ammonia: 1 part of crystals of oxalate of ammonia dissolved in about 24 parts of water, forms a liquor of a

suitable strength. The solution of oxalate of ammonia must not be precipitated or rendered turbid by HS, nor by NH_4S, nor leave any residue after ignition upon a platinum foil.

Oxalate of ammonia is chiefly used to precipitate and detect *lime* in neutral solutions, and to precipitate several metallic oxydes. Its action in this respect is similar to that of oxalic acid, and the reader is accordingly referred to the reactions produced by that acid.

OXALIC ACID.

($\overline{O}$,HO, or C_2O_3,HO) CRYSTALLISED (C_2O_3,3HO.)

Oxalic acid is found in commerce in a state of great purity; but it is also frequently contaminated by nitric acid, by sulphuric acid, and by the presence of organic matter. It can, however, be easily purified by recrystallising it two or three times. It may besides be readily prepared by the operator, by boiling, gently, 1 part of starch with 5 parts of nitric acid, specific gravity 1·42 previously diluted with 10 parts of water, until nitrous acid fumes are no longer evolved, filtering, and evaporating to the crystallising point. Sugar may be used instead of starch; the latter yields about one half, whilst starch yields one eighth only of its weight of crystallised oxalic acid. The crystals should be drained, and then recrystallised.

Oxalic acid should volatilise completely when ignited in a platinum capsule, and should not turn black before its complete volatilisation. It should not turn black either by boiling with sulphuric acid; if it does, it is a sign of the presence of organic matter.

Its crystals should be colourless, and not deliquesce by exposure.

Boiled with water rendered blue by a little sulphate of indigo, the liquid should not be decolourised; if it is, *nitric acid* is present.

Solution of oxalic acid, to which a little HCl has been added, should not be precipitated by BaCl or BaO, NO_5; otherwise, *sulphuric acid* is present.

Oxalic acid should always be kept in crystals, and dissolved only when wanted, since the aqueous solution of this acid in time undergoes decomposition.

The principal use of oxalic acid is for the detection of lime in *neutral* solutions; this last condition is *essential,* since oxalate of lime is soluble in free acids, even, to a certain extent, in free oxalic acid. Baryta, strontia, and a few other earths may also yield precipitates with oxalic acid, but they are soluble in an excess of free oxalic acid, whilst lime is only very sparingly soluble therein, and thorina is quite insoluble in such an excess.

REACTIONS.

Lithia . . .	*Nothing.*	
Baryta		In dilute solutions,
	Nothing .	In very strong, and also in dilute but *neutral* solutions, or with the addition of NH_3,
	White . .	precipitate ($BaO,\bar{O}$), immediately soluble in HCl.
Strontia		After some time,
	Turbidness, or white	precipitate, more rapidly than in solutions of BaO, and rendered abundant and immediate by adding NH_3. The precipitate is immediately soluble in acids.
Lime . . .	*White* . .	precipitate, augmenting by standing, especially with addition of NH_3, very soluble in free HCl, or NO_5, and even in excess of oxalic acid.

Magnesia .	*Nothing* . .	(See Table VII., Observation *x*.)
Alumina . .	*Nothing*.	
Glucina . .	*Nothing*.	
Thorina . .	*White* . .	heavy precipitate, insoluble in excess.
Yttria	*White* . .	In slightly acid solutions, bulky precipitate; soluble in HCl.
Protoxyde of Cerium	*White* . .	precipitate, even in moderately acid solutions; soluble in a great excess of HCl.
Zirconia . .	*White* . .	bulky precipitate; soluble in a great excess of HCl.
Protoxyde of Manganese . .	*White* . .	crystalline precipitate, soluble in acids.
Sesquioxyde of Manganese . .	*Nothing*.	
Oxyde of Zinc	*White* . .	precipitate, in neutral solutions; soluble in acids, in KO, and in NH_3.
Protoxyde of Cobalt	*White* . .	precipitate, with a slight tinge of pink after some time.
Protoxyde of Nickel	. . . . . . *Greenish* .	After some time, precipitate, which augments by standing.
Protoxyde of Iron .	. . . . . . *Yellow* . .	After some time, precipitate, but more rapidly by using a neutral oxalate.
Peroxyde of Iron .	*No precipitate*; but the liquor becomes *yellowish*.	
Oxyde of Cadmium	. . . . . . *Turbidness or white* .	In neutral solutions, immediate precipitate, readily soluble in NH_3.
Protoxyde of Lead .	*White* . .	precipitate, in solutions of neutral salts.
Oxyde of Bismuth	*Nothing* . . *White* . .	at first; after some time, crystalline precipitate.

Suboxyde of Copper	*White* . .	precipitate; becoming greenish-blue by long standing.
Protoxyde of Copper	*Greenish white*	precipitate, in neutral solutions.
Oxyde of Silver .	*White* . .	precipitate; soluble in NH_3.
Suboxyde of Mercury .	*White* . .	precipitate.
Peroxyde of Mercury . .	*White* . .	precipitate; in solutions of corrosive sublimate.
	Nothing.	
Binoxyde of Iridium .	*Nothing.* .	at first; after a time, the solution is *decolourised.*
Peroxyde of Gold .	*Dark greenish-black.*	precipitate (Au); which takes a long time to deposit, except by boiling, CO_2 is disengaged.
Protoxyde of Tin .	*White* . .	precipitate.
Peroxyde of Tin .	*Nothing.* .	even in neutral solutions.
Oxyde of Antimony .	*White* . .	bulky precipitate, which requires a long time to settle, especially if $\bar{T}$,2HO be present.
Titanic acid .	*White* .	precipitate, soluble in acids.
Chromic acid		Reduced to Cr_2O_3, and the liquor becomes
	Greenish, .	especially with the help of heat; CO_2 is evolved.

OXYDE OF BISMUTH (HYDRATED).

(BiO, HO.)

This oxyde is easily prepared by precipitating any salt of bismuth with water, and adding to the milky liquor a very

slight excess of ammonia, heating, separating the precipitate produced by filtration, after it has completely subsided, washing it thoroughly on the filter, and drying it between folds of blotting paper, and finally at a steam heat.

Hydrated oxyde of bismuth is principally used for converting certain metallic sulphurets into their oxydes, by boiling the alkaline solutions of such sulphurets with the hydrated oxyde of bismuth, the latter becoming converted into black sulphuret of bismuth, whilst the sulphuret of the metal originally in solution becomes converted into its respective oxyde. The hydrated oxyde of bismuth is added as long as a black precipitate is produced. It is chiefly employed to convert sulpharsenious and sulpharsenic acids into arsenious and arsenic acids.

OXYDE OF COPPER.

(CuO.)

Protoxyde of copper is obtained in a pure state by igniting the nitrate of that metal. It is used principally as a means of detecting *chlorine*, *bromine*, and *iodine*, but it should then be *entirely free from chlorine.*

The oxyde of copper prepared by the above process is in the state of an exceedingly fine black and very hygrometric powder. The presence of *chlorides* are detected by fusing the suspected compound with microcosmic salt, to which a little oxyde of copper has been added, and directing the flame of the blowpipe upon the bead, when a beautiful blue colour will be perceived round the flame; of course the microcosmic salt used must be quite free from chloride. A more convenient way of testing for chlorides in this way, however, consists in mixing the pulverised compound with microcosmic salt on the hook of a brass or copper wire, and exposing it to the smallest possible flame of a spirit lamp, a beautiful blue flame being then produced if a chloride is present.

Bromides treated in the same way produce also a bluish flame, but which has a tinge of green, and iodides submitted to the same operation produce not a blue but a green flame.

Oxyde of copper is also employed in the elementary analysis of organic substances, and for the detection of organic matter in compounds, because when mixed with them and heated in an appropriate apparatus, the organic matter is burnt, and converted by it into carbonic acid and water. (See Table XXII.—A, General Remarks.)

OXYDE OF LEAD (See PROTOXYDE AND PEROXYDE OF LEAD, AND RED LEAD.)

OXYDE OF MERCURY (See PEROXYDE OF MERCURY.)

PERCHLORIC ACID.

(ClO_7.)

Perchloric acid may be prepared by the method of Dr. Fred. Penny of Glasgow, by throwing chlorate of potash by small portions at a time into hot nitric acid, the result of the reaction being a mixture of nitrate of potash and of perchlorate of potash, which may be readily separated by their difference of solubility. The perchlorate of potash is then decomposed by boiling it with a solution of hydrofluosilicic acid, which yields an almost insoluble salt with the potash, and forming a perchlorate of barytes, which is soluble, and which is subsequently decomposed by a careful addition of sulphuric acid in the usual way, so as to remove all the baryta. Gelatinous silica is then added, in order to remove the excess of hydrofluosilicic acid, fluoride of silicium, which is volatile, being thus formed; it may also be concentrated by distillation, a very weak acid passes off first, but gradually the boiling point rises up to about 400°, and the acid which distils over

then is very concentrated. It is a reagent which is very seldom used, its behaviour with potash, lithia, and ammonia is as follows:—

Potash . . .	*White* . .	abundant precipitate (KO,ClO_7), insoluble in alcohol.
Lithia . . .	*Nothing*. .	in dilute solutions; in very concentrated solutions,
	White . .	precipitate.
Ammonia . .	*Nothing*. .	in dilute solutions; in very strong solutions,
	White . .	precipitate.

PERCHLORIDE OF GOLD (See TERCHLORIDE OF GOLD.)

PERCHLORIDE OF IRON.

(Fe_2Cl_3.)

Perchloride of iron is prepared in the following manner:—Take a certain quantity of clean steel piano-forte wire, and dissolve it in a Florence-flask, with the help of heat, in hydrochloric acid diluted with about six times its bulk of water; decant the clear liquor into a large beaker, add one part more of hydrochloric acid, and, whilst the clear liquor is boiling, add thereto some nitric acid, by small portions at a time, as long as nitrous acid fumes are evolved, or until a small portion of it being pretty largely diluted with water, fails in imparting a blue colour to ferricyanide of potassium test paper, or when tested with solution of ferricyanide of potassium. By the addition of nitric acid, the solution of the protochloride of iron at once becomes of a black or deep brown colour; it is necessary to add the NO_5 only by small portions at a time, and to operate in a capacious beaker, because it frequently happens that the black or brown liquor thus mixed with NO_5

and boiling tranquilly, suddenly undergoes a somewhat violent reaction, accompanied by a copious evolution of nitric oxyde, after which the whole liquor at once resumes its former quiet boiling; it has then become of a deep reddish brown colour, and the whole of the iron is found to be peroxydised, or at least after boiling for a while. The production of the blackish-brown colour above alluded to, is due to a peculiar combination of protoxyde of iron with binoxyde of nitrogen, thus:—

$$9FeCl + NO_5 = 3Fe_2Cl_3 + (3FeO)NO_2.$$

When the whole of the iron is peroxydised, add a slight excess of ammonia to the acid liquor whilst hot; collect the red precipitate, which is peroxyde of iron, on a filter, and wash it thoroughly thereon with hot water. Heat now another portion of diluted hydrochloric acid as above said, and gradually add thereto as much of the well washed peroxyde of iron just alluded to as is sufficient to saturate completely the acid, which is known to be the case when the last portions of peroxyde added refuse to dissolve even by continued boiling. It is important that the solution of perchloride of iron should not contain an excess of acid; this is easily ascertained by filtering the above solution so as to separate the undissolved peroxyde, and testing the clear filtrate by pouring into it a single drop of dilute ammonia; if the small reddish brown precipitate thus produced is not redissolved by shaking the liquor, the solution is all right, if on the contrary the slight precipitate at first produced is redissolved, it is a sign that there is an excess of acid present, and therefore more peroxyde of iron must be boiled with it. When the solution is in a fit condition it is filtered, and the clear brownish yellow filtrate is kept for use.

Perchloride of iron is used principally as a test for phosphoric acid (see Phosphoric Acid) when combined with alkaline earths, and as a means of detecting *succinic* and *benzoic* acids, in which it forms a voluminous reddish-brown precipitate of

benzoate, or of succinate of iron, decomposed by NH_3, hydrated peroxyde of iron being thereby precipitated. It is used also as a reagent for *formic* and *acetic acids*, and also sulphocyanic and meconic acids. (See those acids, and see Table VIII., Observations *j*, *k*, *l*, *m*.)

PERCHLORIDE OF MERCURY.

(HgCl.)

CORROSIVE SUBLIMATE—BICHLORIDE OF MERCURY—PROTOCHLORIDE OF MERCURY.

Perchloride of mercury is found in commerce in an exceedingly pure state; it should volatilise completely when exposed to heat, and be entirely soluble in water, in alcohol, and in ether; it is more soluble in alcohol than in water, and ether takes it up and removes it from both an alcoholic and an aqueous solution.

As a test, 1 part of perchloride is dissolved in about 12 parts of water.

Perchloride of mercury is chiefly used as a test for *phosphorous* and *hypophosphorous acids* and their salts; for *phosphuretted* and *arseniuretted hydrogen*, *hydriodic acid* and *iodides*; and as an oxydising agent, in which case it is converted into subchloride of mercury (Hg_2Cl) or into a grey powder, which is metallic mercury; it is accordingly employed for distinguishing SnO from SnO_2, for if added to a solution of protoxyde of tin a white precipitate of Hg_2Cl is produced, but it frequently happens that the presence of other substances interferes with the use of this reagent.

REACTIONS.

Tetrathionic acid .	*White* . .	precipitate.
Hyposulphurous acid . .	*White* . .	precipitate; not discoloured by time, or ebullition, provided the reagent is in sufficient abundance.

Phosphorous acid	*White* . .	precipitate, insoluble in HCl; augmenting by standing.
Hypophosphorous acid . .	. . . . . .	If the reagent be in excess,
	White . .	precipitate (Hg_2Cl); more abundant than with phosphorous acid and solutions of phosphites.
		If HCl be first added, and then perchloride of mercury poured in, a
	Grey . . .	precipitate of metallic mercury is produced.
Hydriodic acid, Iodides (soluble) .	*Vermilion* .	precipitate.
Boracic acid .	*Reddish-brown*. .	precipitate, insoluble in HO.
Protoxyde of Tin .	*White* . .	precipitate (Hg_2Cl).
Formic acid .	*White* . .	precipitate (Hg_2Cl).

PERCHLORIDE OF PLATINUM. (See Bichloride of.)

PEROXYDE OF LEAD. (See Brown Oxyde of.)

PEROXYDE OF MANGANESE.

(MnO_2.)

Peroxyde of manganese in powder is used in conjunction with red lead, or with brown oxyde of lead, and concentrated sulphuric acid, as a means of detecting the presence of chlorides, an odour of chlorine being then evolved on applying heat. We have already alluded to this reaction in speaking of brown oxyde of lead.

It is used also as a test for *oxalic acid*, a disengagement of carbonic acid taking place when mixed with solution

of that acid or of oxalates, and heated with concentrated sulphuric acid.

MnO_2 serves likewise to detect minute quantities of iodine: for that purpose the compound is mixed with it, and put into a glass tube; the mixture is then moistened with sulphuric acid diluted with its own bulk of water, and heated. Purple fumes of iodine are then immediately given off. (See Table XXVII.—B, Observation *c*.)

Peroxyde of manganese is used also with HCl for generating chlorine.

PEROXYDE OF MERCURY.

(HgO.)

The peroxyde of mercury of commerce is sufficiently pure for analytical purposes, it should, however, be carefully ground into as fine a powder as possible, moistening it first with a little alcohol to prevent the fine powder produced from flying about and being lost. The fine powder thus obtained should then be kept for use in a place sheltered from the sun's rays, for peroxyde of mercury is slowly decomposed by sun light, oxygen being disengaged and the metal reduced.

Peroxyde of mercury is very slightly soluble in water, heated in a test tube it becomes almost black, but it resumes its red colour in cooling, if the temperature be increased to a low red heat, oxygen is evolved, and an abundant sublimate of metallic mercury is formed against the sides of the tube, if the heat be still increased, it may be entirely volatilised and dissipated. It is a powerful oxydising agent. It converts chlorine into hypochlorous acid, sulphurous into sulphuric acid, and chloride of magnesium into magnesia.

Peroxyde of mercury is used in conjunction with potash as a test for hydrocyanic acid, which dissolves it. This is a good test, since peroxyde of mercury is soluble in potash *only in presence of hydrocyanic acid.* (See the tests for HCy in the other Dictionary).

PHOSPHATE OF SODA.

$(2NaO),HO,PhO_5.)$

Neutral Phosphate of Soda.

Phosphate of soda is often contaminated by *sulphate of soda, chloride of sodium, carbonate of soda,* and sometimes even *arsenious* and *arsenic* acids. These impurities are detected as follows:—

To a portion of a dilute solution of the salt, previously acidified with NO_5, add a few drops of a solution of BaCl or of BaO,NO_5, if a white precipitate or turbidness insoluble in water and in acids is produced, it is a sign of the presence of *sulphuric acid* or *sulphate of soda.*

To another portion of the same acidified solution of the salt add a few drops of a solution of nitrate of silver; if this produces a white curdy precipitate insoluble in an excess of NO_5, immediately soluble in NH_3, *chloride of sodium* is present.

If on drenching a portion of the pulverised salt with water and adding a few drops of an acid an effervescence is produced, it is due to the presence of a *carbonate* (carbonate of soda).

If after acidifying slightly the aqueous solution of the salt with *pure* HCl and passing a stream of HS through the liquor, a *yellow precipitate* is produced after some considerable time, especially if the liquid be kept warm, it is *sulphuret of arsenic.*

Pure phosphate of soda fuses at a red heat into a clear limpid glass which is white after cooling.

Phosphate of soda is easily prepared by dissolving a given quantity of commercial phosphoric acid in water, applying heat, and adding carbonate of soda to the hot liquor until a faint alkaline reaction is produced. The liquor is then filtered,

evaporated to the crystallising point, and left at rest. The crystals formed should be redissolved in water, and recrystallised, after which one part of them may be dissolved in about 10 parts of water and the solution kept for use.

Phosphate of soda after exposure to a low red heat becomes converted into PYROPHOSPHATE OF SODA $(2NaO)\ PhO_5$, which is rather less soluble in water than the neutral phosphate, and which produces with AgO,NO_5 a white precipitate $(2AgO)\ PhO_5$, and the liquor is acid after precipitation, whereas with the neutral phosphate a yellow precipitate $(2AgO)HO,PhO_5$ is produced, and the liquor remains neutral after precipitation.

The solution of phosphate of soda to be fit for analytical purposes should not be rendered turbid by NH_3, and the precipitates produced by BaCl or BaO,NO_5 should be immediately soluble in weak NO_5.

Solution of phosphate of soda is principally employed as a test for *magnesia*, for precipitating the '*earths* and certain *heavy oxydes*. Its behaviour is as follows :—

REACTIONS.

Lithia . . .	*No precipitate*,	except by boiling, and then adding ammonia. *Phosphate of potash* produces nothing.
Baryta . . .	*White* . .	precipitate, soluble in free HCl, or in NO_5.
Strontia . .	*White* . .	precipitate, soluble in acids.
Lime . . .	*White* . .	precipitate, soluble in acids.
Magnesia . .	*White* . .	precipitate, in neutral concentrated solutions, especially by boiling, and with the addition of NH_3, or of NH_4O,CO_2. (See Table VI., Observations *m, n.*)
Alumina . .	*White* . .	bulky precipitate, soluble in acids, and in solution of KO.
Glucina . .	*White* . .	bulky precipitate.

Thorina . .	*White* . .	flocculent precipitate.
Yttria . . .	*White* . .	precipitate, in neutral solutions; soluble in HCl, but reprecipitated by boiling.
Protoxyde of Cerium . .	*White* . .	precipitate, in neutral solutions.
Zirconia . .	*White* . .	bulky precipitate.
Protoxyde of Manganese .	*White* . .	precipitate.
Sesquioxyde of Manganese . .	*Brown* . .	precipitate, in exactly neutralised solutions.
Oxyde of Zinc	*White* . .	precipitate, in neutral solutions; soluble in acids, in KO, and in NH.
Protoxyde of Cobalt	*Blue*. . .	precipitate, in neutral solutions.
Protoxyde of Nickel	*White* . .	precipitate, with a green tinge.
Protoxyde of Iron .	*White* . .	precipitate, which becomes *green* by exposure.
Peroxyde of Iron .	*Whitish*. .	precipitate, in neutral solutions, turning *brown* by addition of ammonia; soluble in NH_4O,CO_2 if the phosphate be in excess.
Oxyde of Cadmium	*White* .	precipitate, in neutral solutions.
Protoxyde of Lead .	*White* . .	precipitate; soluble in solution of KO.
Oxyde of Bismuth	*White* . .	precipitate.
Protoxyde of Uranium . .	*Dingy-white*	precipitate, provided too much free acid is not present.

Peroxyde of Uranium . .	*White* . .	precipitate, provided too much free acid is not present.
Suboxyde of Copper . .	*White* . .	precipitate, provided the solution is not too acid.
Protoxyde of Copper . .	*Greenish-white*	precipitate; soluble in NH_3, and the solution is then *blue*.
Oxyde of Silver. .	*Yellow* . .	precipitate $(2AgO)HO,PhO_5$), in neutral solutions; soluble in NH_3.
	White . .	precipitate $(2AgO)PhO_5$), if it be pyrophosphate of soda.
Suboxyde of Mercury . .	*White* . .	precipitate.
Peroxyde of Mercury . .	*White* . .	precipitate.
Oxyde of Palladium . .	*Brown* . .	precipitate.
Binoxyde of Iridium . .	*Nothing*. .	at first, after a time, the solution is *decolorised*, or has only a *greenish tinge*.
Binoxyde of Osmium . .	*Nothing*. .	at first, after a while, a *black* precipitate, and the superincumbent liquor is *bluish*.
Protoxyde of Tin .	*White* . .	precipitate.
Peroxyde of Tin .	*White* . .	precipitate.
Oxyde of Antimony. .	*White* . .	bulky precipitate, requiring a long time to settle.

Protoxyde of Molybdenum	*Brownish-black* . .	precipitate.
Binoxyde of Molybdenum	*Drab* . .	precipitate.
Oxyde of Chromium .	*Light-green* .	precipitate, in neutralised solutions; soluble in excess.
Tellurous acid	*White* . .	precipitate.

PHOSPHATE OF SODA AND AMMONIA. (NaO,NH_4O, HO,PhO_5.) (MICROCOSMIC SALT.)

This compound frequently contains *common salt,* which does not interfere with its use for blowpipe analysis, except when employed with oxyde of copper or with a brass wire for the detection of chlorides, bromides, and iodides (see Oxyde of Copper in this Dictionary, and Chlorides, Iodides, and Bromides in the Dictionary of simple substances, &c.) in such cases, of course, the microcosmic salt must be quite free from chlorides, the presence of which is easily ascertained by dissolving a portion of the microcosmic salt in water, adding nitric acid thereto, and then nitrate of silver; if this produces a white curdy precipitate, insoluble in acids, immediately soluble in NH_3, it is a proof that the salt contains a chloride.

Microcosmic salt should not contain an excess of *phosphate of soda,* and this is ascertained by fusing the salt before the blowpipe upon charcoal, a clear and perfectly transparent bead will then be produced if there be no excess of phosphate,

whereas, in the contrary case, the bead will not be perfectly limpid *after cooling*.

Microcosmic salt is easily prepared by dissolving six parts by weight of crystallised phosphate of soda $(2NaO),HO,PhO_5,26HO$) and 1 part of pure sal-ammoniac (NH_4Cl) in 2 parts of boiling water; as the solution cools, large transparent crystals are deposited, and the mother liquor contains chloride of sodium, which contaminates the crystals: wherefore it is necessary to recrystallise them, after which they may be dried, pulverised, and kept in that state for use.

Microcosmic salt is generally fused on a platinum wire, the hook of which should be small, otherwise it will be found difficult to keep the bead on it.

COLOURS IMPARTED TO A BEAD OF MICROCOSMIC SALT BY SUBSTANCES BOTH ON THE HOOK OF A PLATINUM WIRE AND ON CHARCOAL.

Names of Metallic Oxydes	Oxydising or Outer Flame	Deoxydising or Inner Flame
Baryta . . .		
Strontia . . .		
Lime . . .		
Magnesia . . .		
Glucina . . .		
Yttria . . .	Colourless; but if abundant the bead is cloudy.	Colourless.
Thorina . . .		
Zirconia . . .		
Zinc . . .		
Cadmium . . .		
Peroxyde of Lead		

Names of Metallic Oxydes	Colours of the Bead in the Oxydising or Outer Flame.	Colours of the Bead in the Deoxydising or Inner Flame.
Alumina . .	Colourless, or slight green tinge.	Colourless
Molybdenum . .	ditto.	Green.
Tungsten . .	ditto.	Blue. [If iron is present the bead is red.]
Tellurium . .	ditto.	Grey.
Tantalum . .	ditto.	Grey.
Titanium . .	ditto.	Violet. [If iron is present, the bead is red.]
Bismuth . . .	Colourless when cold; yellow when hot.	Grey.
Chromium . .	Green.	Green.
Copper . . .	Green.	Brown, or reddish.
Antimony . .	Colourless.	Colourless. [If iron is present, red.]
Uranium . . .	Green.	Green.
Lead . . .	Colourless.	Colourless
Cerium. . .	Red when hot; Green when cold.	Colourless
Nickel . . .	Red when hot, greenish when cold.	Red when hot, colourless when cold.
Iron . . .	Red when hot; almost colourless when cold.	Red when hot; colourless when cold.
Manganese . .	Amethyst.	Colourless
Cobalt . . .	Blue.	Blue.
Silver . . .	Yellow.	Grey, or colourless.

Colours of the Bead in the Oxydising or Outer Flame.		Colours of the Bead in the Deoxydising or Inner Flame.	
Colourless	Baryta.	Colourless	Baryta.
"	Strontia.	"	Strontia.
"	Lime.	"	Lime.
"	Magnesia.	"	Magnesia.
"	Glucina.	"	Glucina.
"	Yttria.	"	Yttria.
"	Thorina.	"	Thorina.
"	Zirconia.	"	Zirconia.
	[If in large proportion, the bead is white or cloudy.]	"	Manganese.
		"	Cerium.
"	Alumina.	"	Alumina.
"	Molybdenum.		
"	Tungsten.		
"	Antimony.		
"	Tellurium.		
"	Titanium.		
"	Tantalum.	"	Tantalum.
"	Zinc.	"	Zinc.
"	Cadmium.	"	Cadmium.
"	Lead.	"	Lead.
Green glass	Chromium.	Green	Molybdenum.
"	Uranium.	"	Chromium.
"	Copper.	"	Vanadium.
Yellow	Silver.	"	Uranium.

Colours of the Bead in the Oxydising or Outer Flame		Colours of the Bead in the Deoxydising or Inner Flame.	
,,	Bismuth. Almost colourless after cooling.	,,	Peroxyde of Iron.
,,	Vanadium.		
Red	Cerium.	Red	Tungsten.
,,	Peroxyde of Iron.	,,	Antimony.
,,	Nickel. The tint is faint after cooling.	,,	Titanium.
		,,	Nickel. Becomes faint in cooling.
Blue	Cobalt.	Blue.	Tungsten
Violet	Manganese.	,,	Cobalt.
		Brown, or reddish-brown.	Protoxyde of Copper.
		Violet.	Titanium
		Grey	Tellurium.
		,,	Bismuth.
		,,	Lead.
		,,	Silver.

PHOSPHOROUS ACID.

($PhO_3,3HO.$)

Phosphorous acid is easily and readily prepared by causing a stream of chlorine gas to play over phosphorus kept in a melted state under hot water, the chloride of phosphorus

produced being immediately decomposed as soon as formed into phosphorous and hydrochloric acids, thus :—

$$PhCl_3 + 3HO = PhO_3 + 3HCl.$$

Phosphorous acid is sometimes employed to reduce the salts of gold and of silver. A boiling temperature answers best; in this way solutions of nitrate of silver, of chloride of gold, and of corrosive sublimate, produce a coating of their respective metals.

Phosphatic or *hypophosphoric acid* (Ph_3O_{13}, or $2PhO_5 + PhO_3$) answers the same purpose. This compound is easily obtained by introducing sticks of phosphorus into tubes drawn to a point, but open at both ends placed in a funnel, the neck of which is introduced into a bottle for collecting the acid produced by the slow combustion of the phosphorus. The bottle supporting the funnel and the sticks of phosphorus should be placed under a glass bell, resting upon a flat dish containing some water.

PHOSPHORUS.

Phosphorus is employed sometimes for determining the quantity of oxygen contained in atmospheric air. (See Oxygen in the other Dictionary.)

PICRIC ACID.

$$(C_{12}H_2,(3NO_4,)O,HO.) \text{ or } C_{12}\left\{\begin{matrix} H_2 \\ 3NO_4 \end{matrix}\right\} O,HO.$$

This acid is known also under the name of carbazotic and nitrophenisic acid. There is no difficulty in obtaining it. It is sufficient to heat salicine, or tar oil with 7 or 8 parts of nitric

acid, until ruddy vapours of nitrous acid cease to be evolved. The liquor after such a treatment deposits picric acid on cooling. It may be purified by dissolving it in water and recrystallising.

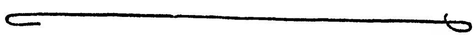
Picric acid, as a reagent, is used only for the detection, or identification of potash, which it precipitates from its solutions and from that of its salts in the form of a pale yellow crystalline powder. If the quantity of potash is small it is best to add alcohol, or to dissolve the salt of potash in that menstruum before testing it, because the carbazotate or picrate of potash is insoluble in it. It must be borne in mind that *ammonia* and *lithia* produce also yellow precipitates with picric acid.

PLATINUM.

(Pt.)

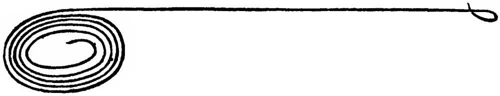
Platinum is used in the state of wire of strips and of foil as supports on which the substances to be examined before the blowpipe are placed. The strip of platinum should be about half-an-inch wide, and about two or three inches long; it is a very convenient support for oxydising substances, in which case the flame of the blowpipe is made to play under the strip of platinum. Of course no substance in the metallic form, or such as are reducible to the metallic state before the blowpipe, should be treated upon the foil, because in such cases the reduced metal will combine with the platinum, and probably fuse or perforate it.

The platinum wire used in blowpipe operations should be as thin as possible, yet not so thin as to bend under the influence of the heat, it is then to be turned at both ends thus—



and it is used by moistening either end and plunging it into the substances to be employed with it, and the portion which

adheres to it is then operated upon. Instead of single pieces of platinum wire, a longer piece may be coiled at one end so as to form a handle, whilst the other end is bent into a loop, thus—



When only short pieces of platinum wire are at hand, one of the ends may be stuck into a piece of cork, and the other end is bent as usual, thus—

These strips or wires should always be kept in bottles or glasses full of water in order that the beads or fused masses may dissolve, and leave the strips or wires clean and always ready for immediate use.

The reagents principally used with the blowpipe are borax, microcosmic salt, carbonate of soda, and cyanide of potassium. These reagents have been treated of each in their respective alphabetical order.

POTASSIUM.

Metallic potassium is now pretty extensively manufactured, and is accordingly readily procurable. It is only used as a test for hydrocyanic acid, or rather to identify cyanide of silver, (see Hydrocyanic Acid in the other Dictionary), also for the detection of phosphoric acid (see that acid in the other Dictionary).

PLUMBIC ACID (PbO_2.) (See BROWN OXYDE OF LEAD.)

POTASH.

(KO.)

Hydrate of potash (KO,HO) is white, of a somewhat crystalline texture, caustic, very alkaline, unctuous to the touch, and evolves a peculiar odour (the odour of lees) when in contact with organic substances. Potash dissolves or decomposes most organic substances, and saponifies fats. It has a great affinity for water, in which it dissolves with a disengagement of heat. It deliquesces by exposure, and becomes gradually converted into its carbonate. It dissolves alumina, silica, and consequently attacks glass and porcelain vessels, wherefore its solution should not be concentrated in vessels made of these materials—capsules or basins of silver only should be used for the purpose. It fuses at a heat below redness, and the crucible or capsule used for fusing it must also be made of silver, for a platinum vessel is attacked by potash, and in fact, by all caustic alkalies in a state of fusion. Pure potash, free from carbonate, is soluble in alcohol.

To prepare caustic potash, dissolve 1 part of *pure* carbonate of potash in about 10 or 12 parts of water in an iron vessel, heat to boiling, and while the solution is thus kept in a boiling state, add a quantity of cream of caustic lime, by small portions at a time, so as not to interrupt the ebullition of the mass, until a small quantity of the liquid, previously filtered, ceases to produce the slightest effervescence with either nitric or muriatic acids. The quantity of lime required is in weight about that of the carbonate of potash employed. When the carbonate of potash has been thus completely decarbonated, the whole mass is left at rest for a few minutes, and the supernatant liquid, previously filtered through a piece of fine linen, is then evaporated in a clean iron vessel,

or better still in a silver capsule, until reduced to about one-third of its bulk (or specific gravity about 1·3); it is then kept for use in small glass bottles provided with stoppers of gutta-percha. If a glass stopper be used, a piece of writing-paper should be first rolled round it, otherwise, as potash dissolves glass, it would in a short time become completely soldered to the neck of the bottle, and its removal rendered impossible.

This aqueous solution is the reagent used in testing. A solution of pure caustic potash may also be obtained by leaving in a bottle the solution of pure carbonate of potash in contact with the necessary quantity of caustic lime in the cold. The mixture is shaken from time to time, and in the course of about twenty-four hours the clear superincumbent liquid may be decanted from the sediment, or siphoned out; it is then fit for use. The carbonate of potash should be anhydrous and pure, and the caustic lime should be freshly prepared.

In order to obtain caustic potash in the solid state, the aqueous solution obtained as just said, is allowed to cool out of the contact of the air, carefully decanted into a capsule of iron, or of silver, and heat is applied until all the water has evaporated and the mass has fused; when in a state of fusion it is poured upon plates of iron, or into cylindrical moulds.

If *pure* carbonate of potash have been employed, the potash obtained is pure also; but if commercial carbonate of potash have been used, the potash produced in that case is always contaminated by *chlorides, sulphates,* and *carbonates.* Pure potash, however, may be prepared from it by treating it with alcohol, which dissolves only the pure potash, and leaves the other salts in an insoluble state—a quantity of spirits of wine equal to about one-third of the weight of the potash is sufficient for the purpose. The alcoholic mixture is stirred and boiled for a few minutes; it is then introduced into a well-stoppered bottle, and left at rest. After a while the liquid will be observed to form three layers; the lower one consists of anhydrous sulphate of potash, and of lime; the

middle one is a solution of sulphate, of carbonate, and of chloride of potassium; the upper layer is the alcoholic solution of pure potash. The latter only should be carefully decanted, or siphoned out, into a retort connected with a Liebig tube-condenser, or into an ordinary still connected with its worm, and about two-thirds of the alcohol is distilled out; the remaining third is then transferred to a silver capsule, rapidly evaporated therein, fused at a dark red heat, and cast upon an iron plate, as before described.

Pure potash may also be obtained by treating crystallised sulphate of potash by hydrate of barytes as follows:—Take about 16 parts of crystallised hydrate of barytes, and dissolve them in boiling water, add now thereto a solution of pure sulphate of potash (about 9 parts of KO,SO_3), until the hydrate of barytes is exactly decomposed, that is to say, until on filtering a small portion of the liquor, acidifying it with HCl and dividing it into two parts, it is found that it is not precipitated by adding solution of KO,SO_3 to the one, nor solution of hydrate of barytes to the other. The whole is then left at rest until the precipitate has completely settled, and the clear superincumbent liquor, which is a solution of pure potash, is kept for use in well-closed bottles, or rapidly evaporated in a silver capsule as mentioned before. The potash, prepared by this process, is preferable to any other for delicate researches, since it is of course completely free from alumina and silica.

Pure solution of potash should be colourless. It is liable to be contaminated by *chloride of potassium, sulphate of potash,* and even *nitrate of potash, phosphates of potash, silica, alumina, lime,* by *carbonate of potash,* and *salts of lead.* These impurities are detected by supersaturating a portion of the potash with pure nitric acid (if the potash is in the solid state, it should of course be first dissolved in water), and tested with *nitrate of silver;* if it produces a white curdy precipitate immediately soluble in NH_3, and reprecipitated by

NO_5, then a *chloride* is present. It is rare, however, to find potash altogether free from chloride of potassium.

If another portion, being tested with *chloride of barium*, yields a white precipitate insoluble in water and in acids, then a *sulphate* is present.

In order to detect the presence of a *nitrate*, take another portion of the aqueous solution of potash, and add cautiously thereto in a test-tube at least one-fourth of its bulk of concentrated sulphuric acid; when the mixture has become cold, drop into it a crystal of protosulphate of iron; if a nitrate is present a deep blackish-brown colour will be observed, especially round the crystal of protosulphate of iron.

Alumina is detected by boiling a portion of the aqueous solution of potash with its own bulk of a solution of sal-ammoniac, in which case a turbidness, or a white flocculent precipitate will appear; or else, and this is a more delicate test, by supersaturating a portion of the solution of potash with HCl, and then adding an excess of NH_4O,CO_2, which will precipitate the alumina.

Silica is detected by supersaturating another portion of the aqueous solution of potash with HCl, and evaporating to perfect dryness. If a gritty, insoluble residue be left after treating the insoluble residue with hot water, it is silica.

The solution of hydrate of potash generally contains a small quantity of *carbonic acid*, which it absorbs from the air; but its presence, except it be abundant, is generally of no moment. Yet only a very slight effervescence should be produced by treating the aqueous solution with an acid, for if the effervescence be brisk, the liquor is unfit for use as a test.

The presence of *lime* is recognised by adding to the aqueous solution a solution of oxalic acid (but less than is necessary to saturate the potash), and boiling. If lime be present a white precipitate of oxalate of lime will appear.

Fused potash is often used for fusing, and thus rendering soluble, substances which, unless so treated, are insoluble in

water and in acids. When used for absorbing gases it need not be quite pure. Potash is used also for the analysis of ammoniacal salts, the ammonia resulting from the decomposition being received in a Liebig's bulb apparatus, charged with a given quantity of sulphuric acid of a known strength; also for the determination of CO_2 and other gases. (See Table XXXIII., Observation *l*.)

It is used likewise for separating from each other certain metallic oxydes which are precipitated at first, but some of which are soluble in an excess of it: such are alumina, oxyde of zinc, oxyde of aluminum, and of lead.

Caustic soda may be obtained in the same manner as caustic potash, and answers nearly all the purposes for which potash is used as a reagent.

The reactions of solution of potash with the various substances are as follow (The precipitates, unless otherwise indicated, are hydrates of the oxydes) :—

REACTIONS.

Baryta . . .	*White* . .	bulky precipitate; almost entirely soluble in a great excess of HO, from which solution, however, it is precipitated again after some time, because the liquor absorbs CO_2 from the air.
Strontia . .	*White* . .	precipitate, same as with baryta.
Lime . . .	*White* . .	gelatinous precipitate; same as with baryta.
Magnesia . .	*White* . .	bulky, flocculent precipitate, insoluble in water. If, however, the liquor contains a sufficient quantity of an ammoniacal salt, *no precipitate whatever* appears until the liquor is boiled for some time, and *provided the potash be in excess.*
Alumina . .	*White* . .	bulky precipitate, in neutral solutions, which precipitate, however, is immediately soluble in an excess of po-

		tash. The precipitate, is reproduced by adding to the alkaline solution one of *sal-ammoniac*, especially by boiling.
Glucina . .	*White* . .	bulky precipitate; soluble in an excess of potash, but reprecipitated by a solution of sal-ammoniac.
Thorina . .	*White* . .	gelatinous precipitate; insoluble in an excess of potash.
Yttria . . .	*White* . .	bulky precipitate; insoluble in an excess of potash.
Protoxyde of Cerium . .	*White* . .	bulky precipitate; insoluble in an excess of potash.
Zirconia . .	*White* . .	precipitate; insoluble in an excess of potash.
Protoxyde of Manganese . .	*White* . .	precipitate, becoming *yellowish*, then *brown*, and finally *black*. If, however, sal-ammoniac is present at the same time, *the precipitate remains white*, but is not so abundant. (See Table V., Observation *p.*)
Sesquioxyde of Manganese . .	*Brown* . .	bulky precipitate. The presence of sal-ammoniac does not interfere with this precipitate.
Oxyde of Zinc . .	*White* . .	gelatinous precipitate; soluble in an excess of potash.
Protoxyde of Cobalt	*Blue* . . .	precipitate, becoming green by exposure, which converts it into peroxyde, insoluble in an excess of potash. A large quantity of sal-ammoniac interferes with the production of this precipitate.
Protoxyde of Nickel	*Apple-green*	precipitate; insoluble in an excess of potash.
Protoxyde of Iron .	*White-greenish*	flocculent precipitate; which becomes *brown* by exposure.
Peroxyde of Iron .	*Reddish-brown*	bulky flocculent precipitate; insoluble in an excess of potash.

Oxyde of Cadmium	*White* . .	precipitate; insoluble in an excess of potash.
Oxyde of Lead . .	*White* . .	precipitate; soluble in a large excess of potash, especially with the help of heat.
Oxyde of Bismuth.	*White* . .	precipitate; insoluble in an excess of potash.
Protoxyde of Uranium . .	*Brown* . .	bulky precipitate; insoluble in an excess of potash.
Peroxyde of Uranium. .	*Yellow* . .	precipitate; insoluble in an excess of potash (KO,U_2O_3).
Suboxyde of Copper . .	*Brownish-yellow* .	precipitate; insoluble in an excess of potash, and gradually by exposure becoming *brownish-black.* The subchloride of copper gives a *white* precipitate, soluble in free HCl.
Protoxyde of Copper . .	*Blue* . . .	voluminous precipitate; becoming *black* by boiling.
Oxyde of Silver .	*Light-brown*	precipitate; insoluble in an excess of potash; soluble in NH_3.
Suboxyde of Mercury . .	*Black* . .	precipitate; insoluble in an excess of potash.
Protoxyde of Mercury . .	*Yellow* . .	precipitate; insoluble in an excess of potash. When in very small quantity, the precipitate is
	Reddish-brown.	If sal-ammoniac is present, the precipitate is
	White.	
Protoxyde of Platinum . .	*Nothing,* .	unless the solution contains some perchloride, in which case a
	Yellow . .	precipitate is produced.
Peroxyde of Platinum . .	*Yellow* . .	precipitate; almost insoluble in free acids; soluble in an excess of KO.

Oxyde of Palladium .	*Brownish-yellow* .	precipitate (subsalt); soluble in an excess of alkali.
Oxyde of Rhodium	*Nothing* . .	except by boiling, when a *yellowish-brown* jelly is formed.
Binoxyde of Iridium . .	*Blue* . . .	colour (like that of copper in ammonia) after heating; but the first addition of the KO either
	Decolourises	the solution, or imparts only a
	Green tinge,	and a
	Slight blackish precipitate . .	is thrown down.
Binoxyde of Osmium .	*Nothing* . .	at first; but by heating, a *black colour,* and by standing a
	Black precipitate .	is produced, the superincumbent liquid is *colourless.*
Peroxyde of Gold .	*Nothing* . .	at first; after a time, the liquor becomes
	Greenish . .	and a slight
	Black . .	precipitate (Au) is produced.
Protoxyde of Tin .	*White* . .	precipitate, soluble in an excess of potash, by boiling, a black precipitate (Sn) is produced, the solution contains a combination of tin and of potash.
Peroxyde of Tin .	*White* . .	bulky precipitate; very soluble in a slight excess of potash.
Oxyde of Antimony .	*White* . .	bulky precipitate; soluble in an excess of potash.
Protoxyde of Molybdenum .	*Brownish-black* . .	precipitate; insoluble in an excess of potash.
Deutoxyde of Molybdenum .	*Brownish-black* . .	voluminous precipitate; insoluble in an excess of potash.
Deutoxyde of Vanadium . .	*Greenish-white*	precipitate; soluble in an excess of potash; and the solution has a *brown colour.*

Oxyde of Chromium	*Light-green*	precipitate; easily soluble in an excess of potash; the whole is precipitated by boiling. (See Table V., Observation *s*; Table XVIII., Observation *h*.)
Tellurous acid	*White* . .	abundant precipitate, in HCl solutions, soluble in an excess of potash.
Titanic acid .	*White* . .	bulky precipitate; insoluble in an excess of potash.
Antimonic acid . .	*White* . .	precipitate, partly soluble in an excess of potash. If organic substances are present, no precipitate is produced.
Antimonious acid	*White* . .	precipitate; partly soluble in an excess of potash. If organic substances are present, *nothing*.
Tartaric acid .	*White* . .	crystalline precipitate, immediately after vigorous shaking.

PROTOCHLORIDE OF MERCURY. (See Perchloride of Mercury.)

PROTOCHLORIDE OF PALLADIUM. (See Chloride of Palladium.)

PROTOCHLORIDE OF TIN.

($SnCl$.)

Protochloride of tin is prepared by boiling in a glass flask a certain quantity of granulated tin with concentrated muriatic acid, taking care to keep the tin in excess, to prevent the production of perchloride of tin. Dilute the solution with water, decant, and keep it in well-closed bottles containing pieces of metallic tin, for without this precaution the solution will absorb oxygen, and rapidly convert the protochloride into a perchloride by transposing with the free hydrochloric acid present.

Protochloride of tin is very easily decomposed, and is, moreover, liable to be contaminated by lead and by iron.

Test it therefore by adding to it a large quantity of hydrosulphuret of ammonia with excess of sulphur, which should completely redissolve, after a time, the dark brown precipitate at first produced; if a *black* precipitate remains in an insoluble state it is probably a sulphuret of lead or of iron.

Solution of protochloride of tin should not be rendered turbid by sulphuric acid. It should immediately produce a white precipitate (calomel) when poured in a solution of corrosive sublimate (perchloride of mercury). Sulphuretted hydrogen should produce a dark brown precipitate, completely soluble, after a time, in a large excess of hydrosulphuret of ammonia, with the help of a gentle heat.

To prepare it, Fresenius recommends the following modus operandi:—

"Fuse a certain amount of English tin in an iron spoon, and after having removed the latter from the fire, triturate the fused mass with a pestle until it has completely solidified. Introduce the powder which is thus obtained into a flask, pour concentrated hydrochloric acid over it (always taking care that the tin predominates), and boil the mixture; dilute the solution subsequently with four times its bulk of water slightly acidified with HCl, and filter. Pour the filtrate into a phial containing small fragments of metallic tin, and close it carefully."

The operator may also simply boil the tin with concentrated HCl, and evaporate the liquor obtained to the crystallising point. As the tin is but slowly taken up by the acid, he may add from time to time a small quantity of NO_5, by which the solution of the tin is greatly expedited. The hydrogen evolved during the process has a very fetid odour.

Protochloride of tin has a great tendency to absorb oxygen and chlorine, by which it becomes converted into bichloride of tin. It is therefore frequently employed as an energetic deoxydising and dechloridising agent. It rapidly and completely reduces a great number of oxydes, such as the oxydes of antimony, of zinc, of mercury, of silver, arsenious and

arsenic acids. It reduces also the peroxydes of copper, of iron, of manganese, molybdic and tungstic acids to lower degrees of oxydisation.

It is chiefly used as a reagent for the detection of gold, in the solutions of which it produces a brown or purple precipitate (purple of Cassius), for which purpose it is advisable to add to the solution of the protochloride a little *cold* nitric acid, in order to peroxydise a portion of the protosalt of tin, for *pure protochloride of tin* produces no precipitate in solutions of gold.

It is also used principally as a test for persalts of mercury, which it converts into subsalts (subchloride of mercury, calomel), and then into metallic mercury. The reactions of protochloride of tin with substances are as follows:—

REACTIONS.

Oxyde of Silver	*White* . .	precipitate, if the quantity of the reagent be small, a larger quantity produces a
	Brown . .	precipitate of metallic silver.
Protoxyde of Platinum . .	*Brownish-red* colour.	
Peroxyde of Platinum . .	*Dark reddish-brown* .	colour (See Table IV., Observation *x*.)
Protoxyde of Palladium. .	*Black* . .	precipitate (Pd); the superincumbent liquor has a beautiful *dark-green colour.*
Peroxyde of Rhodium. .	. . . . . *Dark-brown*	gives to the red solution a colour, but *no precipitate.*
Deutoxyde of Iridium . .	*Light-brown* precipitate.	

Deutoxyde of Osmium .	*Brownish* .	precipitate.
Peroxyde of Gold .	*Purple* . .	colour, and in more concentrated solutions,
	Deep purple or brown	precipitate (purple of Cassius) (Au_2O, SnO_2), ($SnO, SnO_2)4HO$).
Sulphurous acid . .	*Brown colour*, and afterwards	
	Brown or yellowish	precipitate.
Protoxyde of Mercury .	*Grey* . . .	precipitate, which boiling reduces to *globules of mercury*.
Peroxyde of Mercury . .	. . . . . .	a small quantity of reagent
	White . .	precipitate (Hg_2Cl), but if an excess be added,
	Grey . . .	precipitate (metallic mercury).
Tetrathionic acid	*White* . .	precipitate.
Hyposulphurous acid . .	*Brown* . .	precipitate, after a time.
Osmic acid .	*Brown* . .	precipitate; soluble in HCl, and the solution is then *brown*.
Tellurous acid . .	*Black* . .	precipitate, of metallic tellurium.

PROTONITRATE OF COBALT. (See NITRATE OF COBALT.)

PROTONITRATE OF MERCURY. (See NITRATE OF SUBOXYDE OF MERCURY.)

PROTOXYDE OF LEAD. (LITHARGE.)

(PbO.)

The best litharge of commerce reduced to a fine powder is sufficiently pure for the purpose. It is used as a test of the

presence of *acetic acid* or *of acetates*, because when acetates are boiled with dilute sulphuric acid, and an excess of finely pulverised litharge is left to digest with the portion of liquid which distils over, the liquor contains then a basic acetate of lead, which has an alkaline reaction on test-papers. This reaction is quite characteristic, no other acid than acetic acid yielding such a solution. (See Table VIII., Observation *n*.)

PROTOSALTS OF COPPER. (See PROTOSULPHATE OF COPPER.)

PROTOSALTS OF IRON. (See PROTOSULPHATE OF IRON.)

PROTOSALTS OF LEAD. (See NITRATE OF LEAD.)

PROTOSULPHATE OF COPPER.

(CuO,SO_3.)

The protosulphate of copper of commerce, though it often contains small quantities of *sulphate of iron*, and of *sulphate of zinc*, is, nevertheless, pure enough for the purpose to which this reagent is applied; yet it should generally be recrystallised.

Protosulphate of copper in solution is sometimes used as a test for arsenious and arsenic acids. The solution of sulphate of copper should be neutralised by adding as much ammonia as is just sufficient to redissolve the precipitate at first produced, and the solution to be tested should be likewise very nearly neutralised by ammonia, and filtered if necessary. In such a solution protosulphate of copper produces a yellowish-green precipitate with arsenious acid, and a bluish-green precipitate with arsenic acid; both precipitates are immediately redissolved by an excess of either ammonia or of an acid. The test, however, is not characteristic, since solutions

of phosphoric acid and of phosphates, behave in the same manner.

Hydroferrocyanic acid and solutions of ferrocyanides produce with this reagent a brown or a crimson precipitate, which is quite characteristic.

Hyposulphurous acid at first produces *nothing*, but after a time a *turbidness* or a *brown colour* is observed, and by boiling, a *black precipitate* (CuS), falls down, and the superincumbent liquid contains SO_3.

A mixture of 1 part of crystals of sulphate of copper with 2½ parts of crystals of protosulphate of iron, both dissolved in water, is employed to precipitate hydriodic acid as subiodide of copper (Cu_2I), and therefore to separate hydriodic acid or iodides from chlorides and bromides, on which this mixture has no action. (See an example of this method, Table XXI., col. 18, 19.)

It is used also as a test for hydrocyanic acid, a yellowish-green precipitate being produced, which is a mixture of proto and of subcyanide of copper.

Mixed with an excess of caustic potash, it is used for the detection of certain reducing compounds, such as arsenious acid, sugar, &c., in liquids, because in such a case a red precipitate of metallic copper is produced when heat is applied to the mixture.

PROTOSULPHATE OF IRON.

(FeO,SO_3.)

Protosulphate of iron (green copperas) is found in commerce in a state of considerable purity, or at any rate, requires only to be redissolved, filtered, and recrystallised, to be fit for all chemical purposes; yet it is sometimes contaminated by sulphate of copper (blue copperas), sulphate of zinc, or sulphate of magnesia. The preparation of the salt

is, however, so easy, that the operator had perhaps better make it himself. For this purpose he may filter the liquor resulting from the action of sulphuric acid upon protosulphuret of iron in the preparation of sulphuretted hydrogen, and the filtrate will yield on evaporation very pure crystals of protosulphate of iron,—or he may dissolve iron nails free from rust, or pianoforte wire in dilute sulphuric acid, filter, and evaporate the filtrate, which will yield the crystals on cooling. The crystals should be drained, dried, and kept for use in a well-closed bottle, in order to shelter them as much as possible from the action of the air, for protosulphate of iron has a great tendency to absorb oxygen, and it then becomes partly converted into the persulphate of that oxyde. It is on account of this property that it is frequently used as a reducing agent, and therefore as a test for *gold* and *nitric acid*. We have seen in effect that gold is reduced from its solutions by protosulphate of iron in the state of a brown precipitate, or only of a blue colour, according to the state of dilution of the solution. We have seen also that when a solution containing a nitrate is mixed with about half its volume of concentrated sulphuric acid, if a crystal of protosulphate of iron be dropped into the mixture, *after it has cooled*, a dingy greenish-brown colour is observed round the crystal, which colour is due to a solution of nitric oxyde in that of the protosulphate of iron.

Protosulphate of iron is also used as a test for hydrocyanic and hydroferricyanic acids, and ferricyanides, in the presence of which it forms a magnificent blue precipitate (prussian blue). (See these substances in the other Dictionary.)

Protosulphate of iron mixed with protosulphate of copper in the proportion of $2\frac{1}{2}$ parts of the first to 1 part of the second, serves also as a test for iodides and hydriodic acid, a dingy white precipitate being produced. Bromides and chlorides are not affected by this compound reagent.

The solution of protosulphate of iron does not keep, and

should therefore be prepared as wanted by dissolving a little of the crystals in water.

REACTIONS

Oxyde of Silver . .	*White* . .	precipitate; (Ag) in neutral solutions.
Protoxyde of Palladium .	*Black* . .	scanty precipitate in concentrated solutions
	Nothing .	in dilute solutions
Deutoxyde of Iridium	. .	solution is *decolourised*, but
	No precipitate .	is produced.
Peroxyde of Gold . .	*Blue* . .	colour at first, and then a
	Brown . .	precipitate (Au), or if not too dilute, a
	Brown . .	precipitate is at once produced, which is metallic gold. (See Table IV., Observation *u*, Table XVI., Observation *h*.)
Nitric acid .	*Deep blackish brown*	colour This *blackish brown colour* is due to a combination of the protosalt of iron dissolved in binoxyde of nitrogen, by virtue of the following reaction :—

$$10FeO,SO_3 + NO_5 = 3(Fe_2O_3,3SO_3) + (4FeO,SO_3),NO_2$$

Nitrous acid .	*Deep blackish brown*	colour.
Osmic acid .	*Deep black* .	precipitate.
Tellurous acid . .	*Black* . .	precipitate of metallic tellurium.

PROTOSULPHATE AND PERCHLORIDE OF IRON.

$(FeO,SO_3 + Fe_2Cl_3)$

This reagent is exclusively used for the detection of hydrocyanic acid, it is always made *ex tempore* as wanted, by adding some solution of perchloride of iron to one of pro-

tosulphate of the same metal. This reagent, when poured in liquors containing hydrocyanic acid, *to which an alkali has been previously or subsequently added*, yields a blue precipitate (prussian blue), which is quite characteristic.

PROTOSULPHURET OF IRON.

(FeS.)

The best kind of sulphuret of iron is made by putting nails, or iron wire the size of ordinary quills, cut in suitable lengths in a hessian crucible perforated with a hole at the bottom. The hessian crucible charged with the iron should then be placed in a furnace with a good draught, and heated therein to a white heat; if lumps of sulphur be then thrown into the crucible among the white hot pieces of iron, the iron will run off immediately through the hole of the crucible and through the bars of the furnace, in the state of protosulphuret of iron, which may at once be received in a coal shovel placed in the ash-pit.

Protosulphuret of iron is only used for making sulphuretted hydrogen.

PRUSSIATE OF POTASH (YELLOW). (See FERROCYANIDE OF POTASSIUM.)

PRUSSIATE OF POTASH (RED). (See FERRICYANIDE OF POTASSIUM.)

PUCE OXYDE OF LEAD. (See BROWN OXYDE OF LEAD.)

PYROGALLATE OF POTASH.

(Used for the analysis of Oxygen. See other Dictionary, art. Oxygen.)

RED CABBAGE PAPER. (See TEST PAPERS.)

RED LEAD.

(Pb_3O_4, or) $2PbO,PbO_2$)

The composition of this substance is variable. It is decomposed by acids, which precipitate the brown oxyde (plumbic acid PbO_2), and dissolve the protoxyde of lead. This compound is found in commerce in a state of sufficient purity, but it may be easily prepared in the laboratory by Mr. Levol's process, which consists in heating in a clay or hessian crucible, a mixture of 4 parts of litharge, finely pulverised, with 1 part of chlorate of potash. The litharge at first becomes converted into puce oxyde of lead (plumbic acid PbO_2), which may be at once obtained by stopping the operation at that stage, and washing the residue with water. If, however, the operation is continued and the mixture heated to a dark red heat, it becomes thick, oxygen is disengaged, the lead is converted into red lead, which, if the operation be stopped as soon as the signs of decomposition are observed, the mass boiled with caustic potash, in order to remove any protoxyde which may be present, and then with pure water, is left in the state of a crystalline powder of a fine orange red colour, and perfectly pure.

The red lead of commerce is often mixed with *earthy substances, brick-dust, peroxyde of iron*, the presence of which may be detected by boiling the red lead for some time with an aqueous solution of sugar, which dissolves the red lead, but leaves the impurities in an insoluble state.

Red lead is used for the same purposes as the brown oxyde of that metal.

RED LITMUS PAPER. (See Test Papers.)

RED OXYDE OF MERCURY. (See Peroxyde of Mercury.)

RED PRUSSIATE OF POTASH. (See FERRICYANIDE OF POTASSIUM.)

SAL AMMONIAC. (See CHLORIDE OF AMMONIUM.)

SILICA.

(SiO_3)

Silica, or silicic acid, is employed for scarcely any other purpose than with the blowpipe, as a means of testing for *sulphuric acid* or other *sulphur compounds*. The silica used for the purpose should be in the state of an impalpable powder, as obtained from the analysis of silicious minerals. The experiment is made as follows: Fuse upon charcoal before the blowpipe a mixture of silica and of carbonate of soda, so as to obtain a clear, colourless bead, and upon that clear bead place a small portion of the salt under examination; fuse now the whole in the interior flame of the blowpipe, the bead will then become dark brown, or only reddish after cooling if the sulphur is in small quantity in the substance operated upon It is used also as a test for phosphate of of alumina. (See Table XVIII., Observation 2 ‡.)

SILICATE OF POTASH (BASIC).

Silicate of potash (basic), or *liquor silicum*, is a solution of silica in potash, which is made by fusing together 1 part of silica with $2\frac{1}{2}$ parts of carbonate of potash, a residue is left which is soluble in water. This reagent may be prepared also by digesting gelatinous silica in solution of potash. It is only used to detect phosphoric acid in phosphate of alumina; the latter is dissolved in HCl, a solution of pure potash is next added in sufficient quantity to redissolve com-

pletely the precipitate of phosphate of alumina produced, and a solution of basic silicate of potash is then poured in, which precipitates the alumina in the state of a flocculent or gelatinous mass, which is a combination of alumina, potash, and silica; chloride of calcium being now added to the liquor filtered from that precipitate, produces a white precipitate of phosphate of lime, if phosphoric acid is present. (See Table XVIII., Observation *i* ‡)

SODA.

The reactions are the same as potash.

SOLUTION OF INDIGO.

Solution of indigo is made by dissolving in water the indigo paste (sulphate of indigo) of commerce, in such quantity that the solution has only a faint but distinct blue colour. Indigo paste or sulphate of indigo may be prepared by heating 1 part of indigo reduced to powder, with about 10 parts of concentrated sulphuric acid, perfectly free from nitric acid.

Solution of indigo is employed as a test for *nitric acid, chlorine,* and the *nitrates;* in the latter case, however, the nitrate must be decomposed by sulphuric acid, for nitric acid must be in a free state to destroy the blue colour of the solution of indigo.

The reaction with solution of indigo is not alone conclusive, since several other substances (chlorates, bromates, &c.) have the power of producing it.

STARCH.

Starch, as a reagent, is always used in the state of thin, or almost liquid mucilage, which is prepared by triturating common starch with cold water, and heating the mixture to the boiling point, taking care to stir it all the while, to prevent burning. Mucilage of starch, when used as a test for free iodine or bromine, should be quite cold; and as it is acted upon only by free iodine, if the solution to be tested contains metallic iodides, the latter should be decomposed by dropping a little nitric acid in the liquor, if some starch mucilage be then added, the characteristic blue colour of the iodide of starch instantly appears. This blue colour vanishes on heating the liquor, but reappears, though with less intensity, when the liquor has become cold again. Strips of paper, dipped in mucilage of starch and carefully dried, may be used as test-papers for iodine, in the same way as the ordinary test-papers. (See Table XXVI.—B, Observation *b.*)

Starch yields, with bromine, a yellow colour (bromide of starch), which is, however, less characteristic than the blue colour produced with iodine.

SUBNITRATE OF MERCURY. (See Nitrate of Suboxyde of Mercury.)

SUCCINATE OF AMMONIA AND SUCCINIC ACID.

($C_4H_2O_3,HO$ or $\bar{S},HO$.)

Succinate of ammonia is obtained by neutralising a solution of succinic acid with caustic ammonia or its carbonate—but as, in evaporating the solution to the crystallising point, succinate of ammonia parts with a certain quantity of ammonia, the crystals obtained are always *acid*; in order, therefore, to

obtain the solution of *neutral* succinate of ammonia, it is necessary to dissolve the crystals in water and to neutralize the solution with ammonia, or else to dissolve a little succinic acid, and to neutralize the solution with ammonia. It is absolutely necessary that the solution of succinate of ammonia employed should be *perfectly neutral;* for without this precaution the precipitate of persuccinate of iron, for example, which is produced when this reagent is used to separate iron from certain other metals, would partly redissolve in washing. The solution of succinate of ammonia cannot be kept, because it soon turns mouldy When the presence of soda is not objectionable, *succinate of soda* may be used instead of succinate of ammonia. The salt which crystallises from a neutral solution of succinate of soda is a *neutral* and not an *acid* salt.

As the succinates are, or should always be, prepared by the analyst, they cannot but be pure, provided he takes care that the succinic acid employed is pure. Succinic acid crystallises in tables, or in rectangular or three-sided prisms, which are colourless, odourless, somewhat acrid to the taste, entirely soluble in alcohol ; if heated on a platinum foil, succinic acid volatilises completely without alteration , if, on the contrary, it be adulterated with *tartaric acid,* which is sometimes the case, a voluminous residue of charcoal is left on the platinum foil. If it be adulterated by *sulphate* of *binoxalate of potash,* or other *fixed substances,* these bodies will, of course, remain unvolatilised upon the strip of platinum, and their nature may be ascertained, if necessary, by dissolving a portion of the acid in water, and testing with the appropriate reagents ; namely, *chloride of barium* for sulphuric acid, solution of *sulphate of lime* for oxalic acid or an oxalate, &c. Pure succinic acid, triturated with potash, should evolve no odour of ammonia ; otherwise, an ammoniacal salt may be present, perhaps sal ammoniac, or, as I once found it to be the case, *succinate of ammonia.* Succinic acid must be white and free from odour, and its solution, added to one of a

persalt of iron, should not prevent peroxyde of iron from being precipitated, when an excess of ammonia is further added; if, on the contrary, the precipitation of the peroxyde of iron be prevented, it is a proof that succinic acid was mixed with tartaric acid, or some other non-volatile organic substance, such as *citric* or *malic acid;* the admixture of the latter acids, however, is not probable.

Succinate of ammonia is chiefly used to separate *peroxyde of iron* from *protoxyde of manganese,* and *from oxide of zinc;* the solution used for the purpose should be strictly neutral, for otherwise a portion of the precipitated persuccinate of iron will be redissolved, as we just said, by the water used in washing the precipitate—succinate of ammonia is also used sometimes, when hydrofluosilisic acid is not at hand, for distinguishing *baryta* from *strontia* and from *lime.* In effect, neutral succinate of ammonia produces immediately in concentrated solutions of barytic salts a precipitate of succinate of baryta which is soluble in acids. In more dilute solutions the precipitate requires some time to appear, and is partly pulverulent and partly crystalline. In dilute solutions of strontia *no precipitate* is produced, and even in very concentrated solutions some time is required before any precipitate appears; whilst in neutral solutions of lime *no precipitate whatever* is produced by neutral succinate of ammonia, except the solution be very concentrated indeed, and even then a very long time is required before any turbidness takes place.

SUCCINATE OF SODA.

$(NaO,\bar{S}.)$

Succinate of soda is prepared exactly like succinate of ammonia, by neutralising succinic acid with soda instead of ammonia. The crystals which are obtained from a neutral solution of succinate of soda are *always neutral,* whilst we

have seen that the succinate of ammonia which crystallises from a neutral solution, is always an *acid* salt. Succinate of soda answers the same purpose as succinate of ammonia, and may be used in lieu of it when the presence of a fixed salt is not objectionable.

SUGAR.

($C_{12}H_{11}O_{11}$.)

Cane sugar is used as a test of the presence of *free* sulphuric acid, and by this test, which is known under the name of Runge's test, exceedingly small quantities of free sulphuric acid may be rendered apparent. The modus operandi is as follows;—Dissolve 1 part of cane sugar in 30 parts of water, and pour some of it into a small porcelain capsule, or into the cover of a porcelain crucible, and empty it out again, so that the porcelain be only moistened with it, expose it to the jet of steam issuing from a Florence flask containing water in a state of ebullition, and deposit one drop of the suspected liquid in the porcelain capsule or crucible cover just alluded to; if sulphuric acid be present, a stain varying from an intense black colour to a dingy green, according to the quantity of sulphuric acid present, will appear—1 of acid in 8000 parts of water yields yet a very distinct greenish-black stain.

This reaction is due to the decomposition of the sugar by the acid acting upon it at that temperature.

Salts of copper yield also a stain when so treated, but the stain is reddish or yellowish, and cannot be mistaken for that which is produced by sulphuric acid.

SULPHATE OF ALUMINA.

($Al_2O_3, 3\,SO_3$.)

Sulphate of alumina may be prepared by pouring sulphuric

acid on the pure alumina obtained from the analysis of minerals which contain no potash, keeping the mass nearly to a boiling temperature for about two hours, and finally increasing the heat until nearly all the excess of sulphuric acid has evaporated. The mass should then be diluted with a little water, filtered, and the filtrate is kept for use.

Sulphate of alumina is seldom employed as a reagent. A concentrated solution of this substance poured into one of a salt of potash or of ammonia, also concentrated, and previously acidified with muriatic acid, or some other acid, produces a crystalline precipitate of alum (potash or ammonia alum).

It is also used as a test for phosphoric acid, for, by adding sulphate of alumina to a solution containing a phosphate, and adding an excess of NH_3, the whole of the phosphoric acid combines with the alumina, and is precipitated with it. The presence of phosphoric acid is afterwards looked for in that precipitate.

As a test for phosphoric acid a cold saturated solution of ordinary alum answers quite well.

SULPHATE OF AMMONIA.

(NH_4O,SO_3)

Sulphate of ammonia is sometimes used as a test for alumina in the solutions of which it forms a crystalline precipitate of ammoniacal alum, resembling potash alum.

SULPHATE OF COPPER. (See Protosulphate of Copper.)

SULPHATE OF IRON. (See Protosulphate of Iron.)

SULPHATE OF LIME.

(CaO, SO_3.)

Sulphate of lime is prepared as a reagent by pouring dilute sulphuric acid into a concentrated solution of chloride of calcium, the precipitate so produced is then thoroughly washed, the well-washed precipitate is then agitated in a flask with pure distilled water, and the insoluble portion being allowed to settle, the clear superincumbent liquor is decanted into a separate bottle and kept for use. A fresh quantity of water should be agitated with the insoluble portion and left at rest so as always to have a saturated solution in reserve.

Solution of sulphate of lime is used to distinguish *lime* from *baryta* and *strontia;* because, it produces of course no precipitate in solutions of salts of lime, whilst it produces a white one in those of baryta immediately, and in those of strontia after a short time, if, however, the solution of the latter be concentrated, the precipitate is produced immediately, as with baryta. (See Table VI., Observations *c*, *d*. Table XXII.—A, Observation *o*.)

Solution of sulphate of lime is employed also as a test for *oxalic acid* and the *oxalates,* though a few minutes are generally required before the precipitate becomes sufficiently apparent, and as there is no other acid which will form a precipitate with sulphate of lime, or at least so rapidly, this reagent is therefore quite characteristic. *Paratartaric* (racemic acid), however, produces also a precipitate when tested by sulphate of lime, *but not immediately,* a quarter of an hour *at least* is required before the precipitate becomes apparent, whilst with solutions of oxalic acid and of the oxalates one or two minutes are sufficient for the purpose.

SULPHATE OF INDIGO. (See SOLUTION OF INDIGO.)

SULPHATE OF MAGNESIA.

(MgO,SO_3.)

Sulphate of magnesia (Epsom salts) is generally found in commerce in a state of considerable purity. Its aqueous solution, mixed with sal ammoniac and an excess of ammonia, is used as a test for phosphoric acid, especially in solutions which contain sulphuric acid at the same time. The quantity of sal ammoniac should not be more considerable than is necessary to prevent the magnesia from being precipitated by the excess of ammonia.

It is prepared by making a solution of sulphate of magnesia, to which a certain quantity of a solution of NH_4Cl is added, an excess of NH_3 is then poured in; if, as should be the case, a white precipitate appears, it is a proof that the liquor does not contain enough NH_4Cl, which should be now added until the white precipitate is redissolved. The liquor should be well shaken after each addition.

When it is inexpedient or unadvisable to introduce sulphuric acid into the liquor to be examined, a solution of *chloride of magnesium* may be used instead of one of the sulphate, and it answers the same purpose.

Sulphate of magnesia yields also a bulky precipitate, in neutral solutions of phosphates of alkalies especially, by boiling the whole after the addition of the salt of magnesia, provided the solution be pretty concentrated; but the addition of free ammonia increases considerably the delicacy of the test. The mixture should, however, be well stirred and boiled and the precipitate increases by standing. (See Table XIX., Observation *d*. Table XXII.—A, Observation *n*.)

SULPHATE OF POTASH.

(KO,SO_3.)

Sulphate of potash is generally found in commerce in such a state that it is sufficient to dissolve and recrystallise it to obtain it quite pure. Sometimes, however, this salt is contaminated by *sulphate of zinc*, traces of *sulphate of iron* or of *copper, sulphate of magnesia* or *of lime, nitrate* and *bisulphate of potash*.

To detect sulphate of zinc, dissolve a portion of the salt in water, and add an excess of caustic potash, filter if necessary, and test the filtrate with hydrosulphuret of ammonia; if this produces a *white precipitate*, zinc is *certainly* present; if the precipitate is *greyish* or *blackish*, it is due to traces of copper or of iron. If, on adding ammonia phosphate of soda to the original solution, a white precipitate is produced, it is due to the presence of magnesia. *Lime* is recognised by adding oxalate of ammonia, or a soluble oxalate to a portion of the salt dissolved in water, because a *white* precipitate will then be at once produced. If *nitrate of potash* be present, the salt will deflagrate when thrown upon red-hot charcoal; and if it contains any *bisulphate of potash*, it will redden litmus paper.

For use, dissolve 1 part of the pure crystals in about 12 parts of water.

Sulphate of potash as a test is often used instead of dilute sulphuric acid, when it is advisable not to disturb the neutrality of a liquor; for detecting the presence of *baryta* and of *strontia*, and for distinguishing *lime* from baryta and *strontia*, since it produces no precipitate in even concentrated solutions of lime, at least immediately, as may be seen besides from the reactions indicated below.

It is used also as a test for tartaric acid, for which purpose it is preferable to KO, as we have already said.

REACTIONS.

Baryta . . .	*White* . .	immediate precipitate, even in dilute solutions, insoluble in water and in acids.
Strontia . .	*White* . .	immediate precipitate, in concentrated solutions, but only after a few minutes in more dilute solutions.
Lime . . .	*Nothing*. .	even in concentrated solutions, at least immediately.
Thorina . .	*Turbidness*.	provided the solution of the reagent be concentrated, and in excess.
Yttria . . .	*White* . .	precipitate, after some time; entirely, though slowly, soluble, even though the liquor contains sulphate of potash in solution.
Protoxyde of Cerium . .	*White* . .	crystalline precipitate, in moderately concentrated solutions; soluble in water; but insoluble in water which contains sulphate of potash.
Zirconia . .	*White* . .	precipitate (basic salt), after a while; soluble in HCl, except the precipitate has been produced in a hot solution.
Tartaric acid.	*White* . .	crystalline precipitate; soluble in an excess of the reagent.

SULPHATE OF SILVER.

(AgO,SO_3)

Sulphate of silver is prepared by dissolving pure silver in concentrated sulphuric at a boiling heat; the liquor deposits, on cooling, a quantity of small needle-shaped crystals which consist of sulphate of silver. This salt, however, may be obtained in a less troublesome manner by precipitating a concentrated solution of nitrate of silver by one of sulphate of soda; the white precipitate produced, and which is sulphate

of silver, is then washed with cold water. This salt is only very sparingly soluble in water, and less so in cold than in hot water; the latter dissolves only the one-hundredth part of its weight, most of which deposits as the solution cools.

As a reagent, sulphate of silver is never used in qualitative analysis, but in quantitative analysis it is sometimes employed for the determination of the amount of nitrate of soda contained in nitrate of potash.

SULPHITE OF SODA.

($NaO,SO_2,10HO.$)

Sulphite of soda is easily prepared by passing a stream of sulphurous acid gas through a strong solution of carbonate of soda, or in a vessel containing moistened crystals of carbonate of soda, until the evolution of carbonic acid ceases; the mass may then be dissolved in water, and, if acid, neutralised with soda, after which it is evaporated to the crystallising point. The sulphite crystallises in oblique prisms containing 10 equivalents of water, and their reaction is slightly alkaline. The solution of the sulphites of alkalies may serve to identify *permanganic* and *manganic acids*, the solutions of which are *decolourised* when one of a sulphite of alkali is added thereto; and also *osmic* acid, the solution of which, when similarly treated, assumes a *deep violet colour*, and yields a *black precipitate* (Os); after some time the liquor becomes blue, and eventually loses its colour, the precipitate augmenting.

Sulphites are also employed as tests for phosphorous acid, Observation *k*, Table XXI., and as a test for U_2O_3, in the solutions of which it produces a *yellow* precipitate; also for *selenious* acid, the liquor becoming *yellow* and *cinnabar red*, and after a few hours flakes of selenium of a *red colour* are deposited; this reduction, however, is incomplete, when nitric acid is present.

SULPHOCYANIDE OF POTASSIUM.

($KCyS_2$.)

Take 46 parts of ferrocyanide of potassium deprived of its water of crystallisation by exposure to a gentle heat, and mix them intimately with 32 parts of flowers of sulphur, and 17 of carbonate of potash, heat the whole in a crucible until all intumescence has ceased, and a tranquil fusion is obtained; keep the mass for a short time in that state, increasing the temperature to a low red heat. Allow the mass to cool to pasty consistence, and remove it with a spatula from the crucible; when quite cold, break it in pieces, introduce them into a flask, boil them therein with alcohol, and decant carefully, or filter; the liquid on cooling crystallises in long colourless and odourless prisms or in plates of a bitter taste, and which are somewhat deliquescent. The flask containing the remainder of the alcoholic solution or mother liquor should be connected with a condenser and boiled so as to recover the alcohol; the residue in the flask will give again, on cooling, another crop of crystals of sulphocyanide of potassium.

For use, dissolve 1 part of the crystals in about 10 parts of water.

Sulphocyanide of potassium is the most delicate of all tests for persalts of iron, the smallest trace of which is sufficient to produce a more or less intense blood-red colour with that reagent. It is absolutely necessary, however, that the liquor *should not contain a trace of free ammonia*, for the red colour just alluded to would be instantly destroyed by it (peroxyde of iron being precipitated at the same time if enough iron be present), and the reagent would altogether fail to impart the slightest colour to the liquor. On the contrary, the presence of a little free acid does not interfere with or destroy this blood-red colour. The addition of nitric acid in the cold does

not destroy the colour, *except after some time* has elapsed, but if heat be applied, the liquor is instantly decolourised. *Oxalic, iodic, phosphoric,* and *arsenic* acids destroy also that colour, which, however, is reproduced by adding a further quantity of a persalt of iron.

With moderately strong solutions of *protoxyde of copper,* sulphocyanide of potassium yields a *black precipitate,* insoluble in HCl. If the solution is dilute, a *green tinge* only is produced, and a solution of SnCl rendered clear by an addition of HCl produces instantly therein a *white precipitate* of sulphocyanide of copper. It is, therefore, a delicate test of the presence of copper.

In solutions of *nitrate of silver* sulphocyanide of potassium produces a *white* curdy precipitate insoluble in acids and in ammonia.

With *acetate of lead* it produces a *yellow* crystalline precipitate, and with *subacetate of lead* a *white* precipitate.

SULPHURET OF IRON. (See PROTOSULPHURET OF IRON.)

SULPHURET OF POTASSIUM.

(KS)

Sulphuret of potassium is prepared by boiling solution of caustic potash with an excess of sulphur, and decanting the deep yellow liquor produced, mixing it with an equal quantity of solution of caustic KO as at first used. It may also be prepared by passing a stream of HS through a solution of caustic potash to supersaturation, and mixing the saturated solution with as much solution of caustic potash as has been supersaturated with HS.

To obtain the solution of KS with excess of sulphur, do not

add the second portion of caustic potash, but take the liquor which has been supersaturated with IIS, or boiled with the sulphur.

Sulphuret of potassium is used instead of hydrosulphuret of ammonia, to separate copper in the state of sulphuret, because sulphuret of copper is quite insoluble in sulphuret of potassium, whilst, on the contrary, it is soluble to a certain extent in hydrosulphuret of ammonia. (See Table XXIII., Observation *s*.)

SULPHURET OF SODIUM.

(NaS)

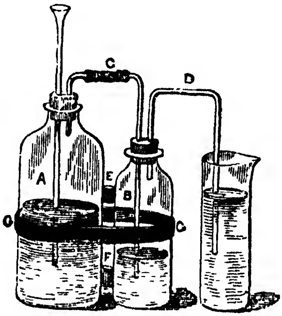
Sulphuret of sodium is prepared precisely like sulphuret of potassium, but it may be obtained in the solid state by passing a stream of IIS through a strong solution of caustic soda, which absorbs the gas instantly, no matter how rapid the disengagement may be; and the disengagement tube plunging into the solution should be of a large bore that it may not become obstructed by the crystals of sulphuret of sodium, which are eventually formed when the soda solution is sufficiently strong. The solution of sulphuret of sodium may be kept for a long time without alteration; it acts like sulphuret of potassium, and may often be used with advantage instead of both sulphuret of potassium and hydrosulphuret of ammonia.

The crystals of sulphuret of sodium which are produced in concentrated caustic soda solutions treated by HS should be drained and redissolved in hot water; as the liquor cools, large prismatic, colourless, and transparent crystals of sulphuret of sodium are then deposited, which should be kept in well-stoppered bottles.

SULPHURETTED HYDROGEN.

(HS.)

Sulphuretted hydrogen is generally prepared in the laboratory by treating protosulphuret of iron by dilute sulphuric acid, or hydrochloric acid. (See Protosulphuret of Iron). A few lumps of the protosulphuret are placed in a bottle A, and covered with water, concentrated sulphuric acid is poured



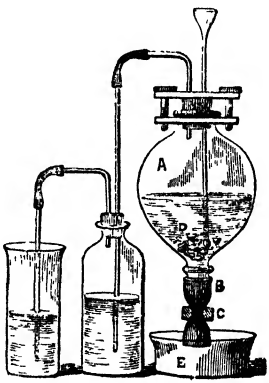
through the funnel tube B, this immediately produces an evolution of sulphuretted hydrogen gas, which passes through a washing bottle C, containing some water for the purpose of purifying the gas, for when the current of the gas disengaged is rapid it always carries off with it small quantities of sulphuric acid and of protosulphuret of iron, which it is essential to arrest in the washing bottle C, before allowing it to pass into the beaker D, which contains the solution which has to be submitted to its action. E E′ are connections of vulcanised India-rubber. The disengagement bottle should be emptied,

and the lumps of protosulphuret of iron as yet undissolved, and likewise the disengagement tube F should be well washed, in order that everything should be in readiness for any future operation. When, however, the operator has not frequent occasion to use the apparatus, he should not only wash the lumps of protosulphuret of iron left in the bottle, but dry them rapidly by passing them through warm sand, for otherwise they will become coated over with peroxyde of iron, which will render them almost unattackable by the acid, and prove a source of vexation and of loss of time. When this happens, the operator should pour into the bottle containing the protosulphuret of iron, and the sulphuric acid and water, a certain quantity of strong muriatic acid, which will promote the solution of the peroxyde of iron, and the disengagement will then proceed as usual.

The most convenient apparatus, perhaps, for the purpose, if I may be permitted to say so, is that which I contrived about two years ago, and by means of which the emptying and cleaning of the bottles is greatly facilitated; whilst, on the other hand, the lumps of protosulphuret of iron remaining under ordinary circumstances in an atmosphere containing some HS, are less exposed to peroxydisation, and need not be removed, so that the apparatus is always in a state fit for immediate service. I am aware that several other apparatus have been contrived for the purpose of keeping a supply of sulphuretted hydrogen in constant readiness, but I think that they generally have many defects, of which, I believe, that mine is entirely free, or else they are of such a size as to be unsuited to a private laboratory. To sum up in a few words and in a collective manner the objections to those apparatus, I may say, in the first instance, that they are expensive, and that being provided with glass stoppers, or, still worse, with glass cocks ground in, they always become fixed in their seats, often permanently so, either by the adhesion of the two surfaces, or by the interposition of crystals of protosulphate of

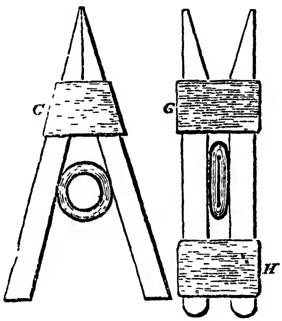
iron, resulting from the action of the dilute sulphuric acid upon the protosulphuret of iron employed; that the sulphuret of iron remaining constantly immersed to a greater or less extent in the acid water, either continues uselessly to generate sulphuretted hydrogen so as to constitute a machine perpetually evolving an offensive odour, or in a few days becomes covered with crystals of sulphate of iron, or enfilmed with peroxyde of iron, either of which conditions impedes the action of the acid, and consequently the production of sulphuretted hydrogen when next wanted.

The apparatus represented below, on the contrary, is inexpensive, is easily managed, is always ready to yield an abundant supply of gas, the sulphuret of iron employed never needs remain in contact with the acid longer than is necessary, nor, under ordinary circumstances, does it become coated with peroxyde of iron, because it remains immersed in an atmosphere of the gas; nor with crystals of green copperas, because it is never left in contact with a saturated liquor.



The apparatus consists of a bottle, A, open at both ends, and provided at the upper extremity with a sound and well-fitting

cork, perforated with two holes, through one of which a glass funnel tube, B, passes, whilst the other hole is provided with a bent tube, C, connected by means of a caoutchouc tube with a straight piece of glass tube D plunging into a washing bottle E, from which a disengagement tube F issues in the usual way. The lower extremity of bottle A is provided with a well fitting vulcanised India-rubber tube of suitable bore, which may be closed by means of a strong squeezer or clamp. The operator, however, can easily make himself a clamp equally effective, though not so good looking, by placing the tube between two pieces of wood tied



at each end by strong vulcanised Indian-rubber bands, G, H, so as to keep the tube closed, and removing one of them, namely, the band H, when it is desired to open it, as represented in the figure. To use the apparatus, proceed as follows:— Remove the cork of bottle A, and introduce into the latter a certain quantity of pretty large lumps of protosulphuret of iron, so as to fill up the lower extremity of the bottle; return the cork to its place; pour now through the glass funnel tube B as much water as is sufficient to cover the lumps and to close the lower extremity of that funnel tube, and then a suffi-

cient quantity of sulphuric acid as usual When the experiment is at an end, unscrew the clamp H, the liquid in bottle A at once falls down through the Indian-rubber tube G into a beaker placed underneath, and may be used for a fresh operation if not spent or saturated, the clamp H is then screwed down again and fresh water is poured through the tube funnel, so that when next wanted, the operator has only to pour concentrated sulphuric acid into the apparatus as before to obtain at once an abundant disengagement of sulphuretted hydrogen.

In a laboratory where several operators are engaged, the washing bottle may be of larger dimensions, and have other disengagement tubes provided with small clamps, so that several experiments may be carried on with the apparatus at the same time.

Of course, all the joints of the bottles must be perfectly air-tight.

Sulphuric acid may be replaced by hydrochloric acid;

And protosulphuret of iron and sulphuric or hydrochloric acid may be replaced by sulphuret of antimony and hydrochloric acid, but in that process, as it is necessary to apply heat, the apparatus last described cannot be used.

Aqueous solution of sulphuretted hydrogen is prepared by passing a current of sulphuretted hydrogen gas through cold distilled water which has been previously well boiled in order to expel the atmospheric air, which has a decomposing action upon that solution. The solution should be quite saturated, and this is known by closing the mouth of the bottle with the thumb and strongly shaking the contents. If the thumb be gently pushed off, the water is saturated, if, on the contrary it is sucked in, the current must be continued. When a liquid has to be tested by sulphuretted hydrogen, a beaker or bottle containing the liquid under examination is substituted and the gas is passed through it. Solution of sulphuretted

hydrogen in water should be kept in small bottles quite full and with their mouths inverted in water, thus:

Aqueous solution of sulphuretted hydrogen is not fit for use unless it smells strongly of the gas, is perfectly clear, is not turned black by ammonia, and yields an abundant white precipitate of sulphur when tested by perchloride of iron.

When the metal whose presence is sought for in a liquid is one which can form with sulphuretted hydrogen, a sulphuret of a dark colour, the operator may put a small piece of proto-sulphuret of iron into a test tube, pour upon it moderately diluted sulphuric acid, and hold in the tube a piece of paper moistened with the liquid under examination.

Sulphuretted hydrogen is one of the most important reagents, not only as a means of identifying and separating certain metals, but, as we have already seen in the other Dictionary, under the head Hydrosulphuric Acid, as a means of classing the metallic oxydes into distinct groups; for we have seen that when sulphuretted hydrogen gas is passed through the *acid* solutions of certain metallic oxydes a transposition takes place, the oxygen of the oxyde forming water with the hydrogen of the gas, and the sulphur thereof combining with the metal so as to produce an insoluble sulphuret, which separates. Certain other metallic oxydes, however, not being thus precipitated from their acid solutions, but only from their neutral solutions, or after the free acid has been neutralised, it follows that by adding an acid to a solution of several metallic salts, one group of them only will be precipitated, whilst another group will

remain in solution. (See Hydrosulphuric Acid in the other Dictionary.)

Sulphuretted hydrogen is also employed as a reducing agent; and accordingly the persalts of iron, and chromic acid, for example, are converted by it into protosalts of iron, and sesquioxyde of chromium; in such cases the sulphur of the sulphuretted hydrogen always separates in the state of a white precipitate.

The action of sulphuretted hydrogen with various substances is as follows: (See Table IV., Observation *k*.)

REACTIONS.

Protoxyde of Manganese	*Nothing*	in neutral or in acid solutions, except the acid is a weak one (such as acetic acid, &c), but if ammonia be then added, a
	Flesh-red .	precipitate is produced.
Sesquioxyde of Manganese	*Milk-white*	precipitate of sulphur, and reduction to the state of protoxyde. (See Table IX, Observation *k*.)
Oxyde of Zinc	*Nothing* . .	in acid solutions; but if the acid is a weak one, or in alkaline solutions,
	White . .	precipitate (ZnS).
Oxyde of Cobalt	*Nothing* . .	in acid solutions, except the acid be a very weak one, in which case the liquor becomes *blackish*, and eventually a *scanty-black* precipitate (CoS) falls down, but *if the solution is alkaline* a
	Black .	precipitate is at once produced.
Oxyde of Nickel	*Nothing* . .	in acid solutions; but if the solution is alkaline, or the acid is a very weak one, a
	Black . .	precipitate is at once produced.
Protoxyde of Iron	*Nothing* . .	in acid solutions, except the acid is a weak one; in which case a

	Black . .	precipitate is produced. Alkaline solutions yield at once a
	Black . .	precipitate, or the solution becomes merely *greenish-black*, if only a trace of iron is present.
Peroxyde of Iron .	*Milk-white* .	precipitate of sulphur, in acid solutions, and the peroxyde is reduced to the state of protoxyde, which remains in solution. (See Table IV., Observation *a'*, Table IX., Observation *k*.)
Oxyde of Cadmium .	*Yellow* . .	precipitate, in alkaline and in acid solutions.
Protoxyde of Lead .	*Black* . .	precipitate, in neutral, alkaline, or acid solutions. If only a trace is present, the liquor assumes a brown colour.
Oxyde of Bismuth.	*Black* . .	precipitate, in acid, alkaline, or neutral solutions, if only a trace be present, the liquor assumes merely a *brown* colour.
Peroxyde of Uranium .	*White* . .	precipitate (sulphur), and the salt is reduced to a protosalt.
Suboxyde of Copper . .	*Brown* . .	precipitate, in acid, alkaline, or neutral solutions.
Protoxyde of Copper . .	*Black* . .	precipitate, in neutral, acid, or alkaline solutions.
	Dark-brown	precipitate; if only a small quantity be present.
Oxyde of Silver	*Black* . .	precipitate, in acid, alkaline, or in neutral solutions.
Suboxyde of Mercury . .	*Black* . .	precipitate, in acid, alkaline, or neutral solutions. The precipitate is insoluble in the dilute acids (see Table XVII., Observation *a*), soluble in aqua regia. The same observation

		applies to peroxyde of mercury. (See Table XXIII, Observation *s*.)
Peroxyde of Mercury . .	. . . . . . *White* . .	In acid, alkaline, or neutral solutions, precipitate, at first, becoming
	Yellow . .	then
	Reddish . .	and finally
	Black.	(See Table IV, Observation *t*)
Protoxyde of Platinum . .	*Brown* . .	colour, at first in acid or neutral solutions, but after some time
	Black . .	precipitate.
Peroxyde of Platinum . .	*Brown* . .	colour at first, in acid or neutral solutions, but afterwards
	Black . .	precipitate.
Protoxyde of Palladium . .	*Black* . .	precipitate, in acid, neutral, or alkaline solutions.
Peroxyde of Rhodium . .	*Brown* . .	precipitate, in acid, alkaline, or neutral solutions, but the superincumbent liquor is *brown*
Binoxyde of Iridium . .	*Brown* . .	precipitate, after a time, in acid or neutral solutions, but the solution is at once *decolourised.*
Binoxyde of Osmium .	*Brownish yellow* .	precipitate, after a time, in acid, alkaline, or neutral solutions.
Peroxyde of Gold .	*Black* . .	precipitate, in acid or neutral solutions; the precipitate is insoluble in the simple acids; soluble in a great excess of NH_4S.
Protoxyde of Tin .	*Dark-brown*	precipitate, in acid or neutral solutions. (See Table XV., Observation *h*; Table XVI., Observation *d*, Table XXIII., Observation *q*.)
Peroxyde of Tin .	*Yellow* . .	precipitate, after a time, augmenting by standing. (See Table IV., Observation *c'*, Table XVI., Observation *d*)

Oxyde of Antimony	*Orange-red*	precipitate, in acid or neutral solutions; if T2HO is present, an orange-red colour only is produced. (See Table XVI., Observation *f*; Table XXIII., Observation *o*)
Protoxyde of Molybdenum	*Nothing* .	at first, in acid and neutral solutions after a time
	Brownish-black .	precipitate.
Deutoxyde of Molybdenum	*Nothing* .	at first, in acid and in neutral solutions, but after a time
	Brown . .	precipitate.
Arsenic acid .	*Yellow* . .	precipitate, especially in acid solutions, and with the help of heat, soluble in NH_3 and in NH_4S (See Table IV., Observation *c′*; Table XVI., Observation *c*, Table XXIII., Observation *g*.)
Arsenious acid	*Yellow* . .	precipitate, especially in acid solutions, and by warming, soluble in NH_3 and NH_4S. (See Table XVI, Observation *c*.)
Sulphurous acid . .	*Milk-white*.	precipitate of sulphur. In solutions of sulphites an acid must first be added
Selenious acid . .	*Lemon-yellow*	precipitate (SeS).
Tellurous acid . .	*Dark-brown*	precipitate; very soluble in NH_4S.
Telluric acid .	*Light-brown*	colour, after a time, and a metallic coating (TeS) takes place.
Chloric acid .	*White* . .	precipitate (sulphur).
Iodic acid . .	*Brown* . .	precipitate (iodine); soluble in a large excess of aqueous solution of HS with deposit of sulphur.
Antimonic acid . .	*Orange-yellow* .	precipitate in acid solutions; soluble in NH_4S, reprecipitated by HCl.
Antimonious acid	*Orange-yellow*	precipitate in acid solutions.

Molybdic acid	*Brown*	precipitate; the superincumbent liquor is *green.*
Vanadic acid	*Greyish-brown*	precipitate in acid solutions (mixture of oxyde of vanadium and of sulphur.)
Chromic acid	*Green*	At first the liquor becomes *Green* and there is a deposit of sulphur, and the liquor contains SO_3 in solution. (See Table XV., Observation *e*; Table XXIII, Observation *h*.)
Permanganic acid	*Decolourised*	and deposit of sulphur.
Manganic acid	*Decolourised*	and deposit of sulphur.
Osmic acid	*Brownish-black*	precipitate; easily deposited, by adding HCl or other free acid.
Bromic acid, Bromates	*White*	precipitate (sulphur); and the acids are reduced to the state of hydrobromic acid, or of bromides, with formation of SO_3.

SULPHURIC ACID (SO_3).

OIL OF VITRIOL.

(SO_3HO)

The sulphuric acid, or oil of vitriol of commerce, is seldom pure, but generally contains *sulphate of lead,* owing to which it becomes milky, and deposits a white sediment when diluted with water, because sulphate of lead is soluble in concentrated sulphuric acid, but is insoluble, or only sparingly soluble, in the dilute acid.

Nitric acid also is frequently present; this impurity is detected by boiling a portion of it with water tinged blue with sulphate of indigo, in which case the blue colour will disappear; also by dropping a crystal of protosulphate of iron

into the concentrated acid, in which case a brownish colour all round the crystal will appear; and lastly, by slightly diluting it with water and boiling it with a few copper filings, in which case nitrous acid fumes will be evolved.

Arsenious acid is another frequent impurity, which is detected by diluting the acid with water, and passing through it a stream of sulphuretted hydrogen, which, in that case, will produce, after a time, a *yellow* precipitate. If the precipitate is *black*, it is due to the presence of *lead*; if *brown*, to the presence of tin, or traces of *lead*.

The most convenient way, perhaps, of testing for arsenic consists in diluting the acid with about twice its bulk of water, and pouring it over some *pure zinc* (see Zinc) contained in a bottle provided with a disengagement tube, which, after a few moments, is brought to a red heat by holding a spirit lamp under it; if arsenic be present, a metallic mirror of arsenic will be deposited in the cold part of the tube. (See the tests for Arsenic.)

Sulphuric acid should not leave the slightest residue by evaporation; if it does, it is probably due to the presence of sulphate of potash, or of lead, by which the commercial acid is occasionally contaminated.

All these impurities may be removed by distillation, but when only a dilute acid is wanted, and lead only is present, it may be completely separated by adding water, and allowing the white precipitate of sulphate of lead to subside. The clear superincumbent acid is then fit for use.

If the acid contains arsenic, it must of course be purified before it can be used with Marsh's apparatus. The most ready way, perhaps, of removing arsenic consists in diluting the acid with five or six times its bulk of water, supersaturating the mixture with sulphuretted hydrogen, and allowing the acid liquor to remain at rest for about twelve hours in a warm place. The yellow precipitate of sulphuret of arsenic, which will then have settled, may be removed by filtering, and

the filtrate is then boiled until all odour of HS has disappeared.

The ordinary commercial sulphuric acid, however, may be used for most analytical operations without purification. But when concentrated and pure sulphuric acid is required, the concentrated commercial acid must be distilled, the first portions which come over being rejected.

Sulphuric acid, especially in the concentrated state, and with the help of heat, is a powerfully oxydising agent, very few substances resisting its action; yet it is seldom employed as a solvent, because its elimination requires a very high temperature.

The concentrated acid is chiefly used (with alcohol or methylic spirit) as a test for *boracic acid*, the alcohol burning then with a beautiful green colour when inflamed; it is used also for the purpose of liberating most other acids, and also *bromine*, *iodine*, and *fluorine*, from their combinations.

Concentrated SO_3,HO, in the cold, is also used as a test for the presence of nitrates, bromates, chlorates, and iodates, as shown, Table I., A., col. 3 to 6, and to distinguish them from each other. If the base of these salts is baryta or strontia, some time must elapse before the reaction is produced.

Concentrated sulphuric acid, with the help of heat, is also used as a test of the presence of oxalates, &c., (see Table I. —C., Observation *g*,) and, in fact, as a preliminary test for all acids, as shown in Table I., C.

It is also frequently employed for destroying the organic matter mixed with or existing in certain compounds, and which might interfere with the analysis thereof. (See Table I., D., and Observations.)

Also as a test for formic acid. (See Table XXVII.—B., Observation *b*.)

For iodine and iodides. (See Table XXVII.—B., Observation *c*.)

Concentrated sulphuric acid is used also to absorb moisture from powders, precipitates, and gases.

In the dilute state it is used as a special reagent for *oxyde of lead, baryta, strontia,* and *lime*; instead of the dilute acid, the solution of a sulphate is sometimes preferable as a test.

It is used also to obtain hydrogen and sulphuretted hydrogen.

The reactions of sulphuric acid and of soluble sulphates upon substances are as follows:—

REACTIONS.

Baryta . . .	*White* . .	precipitate; insoluble in water and in acids; very sparingly soluble in NO_5.
Strontia . .	*White* . .	precipitate; insoluble in the dilute free acids.
Lime . . .	*White* . .	bulky precipitate, in concentrated solutions, and
	Nothing . .	in dilute solutions.
Alumina . .	. . . .	with an addition of KO,CO_2; and if the solution is not too dilute,
	Crystalline .	precipitate (alum).
Thorina . .	*Turbidness*	and precipitate by KO,SO_3, in concentrated solutions. It is a potash sulphate of thorina.
Yttria		(By KO,SO_3),
	White .	precipitate (double sulphate of yttria and potash), sparingly soluble in water.
Protoxyde of Cerium . .	. . . . .	(By KO,SO_3), if the solution is not too dilute,
	Crystalline	precipitate, immediately, or after a short time.
Zirconia		(By KO,SO_3),
	White . .	precipitate, after some time (double sulphate); soluble in a great excess of HCl.

Protoxyde of Lead .	*White* . .	precipitate; almost insoluble in water. In solutions which contain $\overline{\overline{T}}$,2HO dilute SO_3 produces nothing, provided the quantity of lead be very small, otherwise the presence of organic matter does not interfere. Sulphate of lead is decomposed by HCl, and is converted thereby into PbCl. (See also Table XIII., Observation *d*)
Oxyde of Antimony .	*White* . .	precipitate; insoluble in excess, and the filtrate contains still a great deal of the oxyde.
Perchloric acid	. . . . . *Yellow* . .	The mixture with SO_3HO becomes and by boiling chlorine and oxygen are disengaged.
Chloric acid	. *Greenish-yellow gas*	SO_3,HO poured *in the cold* on a chlorate, produces a (hypochloric acid ClO), the odour of which is characteristic, and resembles chlorine. The first impression of the acid is to impart to the dry salt a *deep yellow* colour. Operate only on very small quantities, and eschew heat for fear of explosion (See Table VII., Observation *v'*.)
Bromic acid .	*Hyacinth-red gas* (Br).	
Hydrobromic acid Bromides	. . . . . . *Hyacinth-red*	SO_3,HO, and heat applied, evolve fumes (Br), and SO_2 is also disengaged.
Boracic acid, Borates	*Crystalline spangles* *Green flame*	If mixed with SO_3,HO, and methylic spirit, or with alcohol, and inflamed, (See Table XXVI.—A., Observation *c*.)
Titaniates of alkalies . .	*White* . .	precipitate; soluble in an excess of the reagent.
Tungstates of alkalies . .	*White* . . *Yellow.*	precipitate, becoming
Permanganic acid .	*Brown* . .	precipitate (MnO_2), with the help of heat.

Hydriodic acid, Iodides .	. . . . . *Violet fumes*	If SO_3HO be poured upon the compound and heat applied, (See Table XXVII.—A, Observation *b*)
Hydrofluoric acid, Fluorides .	. . . . *Corrodes glass.*	Treated by SO_3,HO,
Oxalic acid	*Blue flame* .	Treated by SO_3,HO, and heat being applied, a gas is evolved (CO), which burns with a on being ignited. (See Table VII., Observation *x*.)
Citric acid	*Blue flame* .	Frothiness and disengagement of CO_2 and CO, which burns with a especially after having absorbed the CO_2 with a little KO.

SULPHUROUS ACID.

(SO_2)

Sulphurous acid is prepared in a very economical manner by filling about half the capacity of a florence flask with concentrated sulphuric acid, and then adding a pretty large quantity of charcoal in small lumps, heat is then cautiously applied, and shortly afterwards a gaseous mixture of sulphurous acid, of carbonic acid, and of oxyde of carbon is evolved, the gases being made to traverse a vessel filled with water and kept as cold as possible until it is saturated. The aqueous solution of sulphurous acid should be kept in well stoppered bottles.

Sulphurous acid may also be prepared by means of metallic copper or mercury and concentrated sulphuric acid. The acid and the metal are introduced into a florence flask, and heat is at first cautiously applied; as soon as the first bubbles of sulphurous acid begin to be disengaged, the source of heat must immediately be removed, for otherwise the gas would be

disengaged with almost uncontrollable violence. The action of mercury, however, is much less violent than that of copper The gas is received, as before, in bottles full of water as cold as possible, and well stoppered.

Alcohol being capable of dissolving much more sulphurous acid than water, is sometimes substituted for that liquid.

The solution of sulphurous acid in water absorbs oxygen very rapidly, and becomes thus partly converted into sulphuric acid, which, however, does not interfere with its use as a reagent. This tendency of sulphurous acid to absorb oxygen renders it available as a reducing agent, and for converting higher degrees of oxydisation into lower ones; it is thus that it precipitates metallic mercury or gold from their solutions; and thus it reduces chromic and arsenic acids to the state of oxyde of chromium and of arsenious acid. Yet, for these purposes a sulphite of alkali added in the solid state, or dissolved at the time, and a few minutes before using, is often preferable.

Sulphurous acid is employed also to precipitate selenious and tellurous acids from their solutions.

The aqueous solution of sulphurous acid should always have a strong odour of the gas; when this is not the case, it is no longer fit for use.

The behaviour of certain substances with sulphurous acid is as follows:—

REACTIONS

Salts of Silver	*White* . .	precipitate, which, by boiling is converted into metallic silver, which coats the vessel with a bright metallic film.
Protosalts of Copper	*Light brown*	bulky precipitate.
Perchloride of Gold .	*Metallic Gold* . .	reduced in the cold.
Telluric acid .	*Black* . .	pulverulent precipitate.

Tellurous acid	*Black* . .	precipitate; especially by boiling.
Selenious acid	*Cinnabar red*	precipitate of selenium (in the cold.) (See Table XXXII., Observation *b*.)
Protochloride of Tin . .	*Brown* . .	precipitate (SnS).
Protosalts of Lead .	*White* . .	precipitate; soluble in cold nitric acid.
Chloride of Barium .	*White* . .	precipitate (in neutral solution); soluble in the dilute acids.
Chloride of Calcium .	*White* . .	precipitate; soluble in the dilute acids.
Chloric acid		converts sulphurous into sulphuric acid.
Permanganic, and Manganic acids	*Decolourised*	instantly.
Osmic acid .	*Deep violet blue* . .	tinge (?). (See Table XVII., Observation *h*.)
Phosphorous acid	*Milkiness* .	(See Table XXI., Observation *k*.)
Chromic acid	*White precipitate or Milkiness*	(sulphur) the acid is reduced to the state of Cr_2O_3 the liquor becoming of a
	Greenish colour	and contains then SO_2 and S_2O_5.

TARTARIC ACID.

($C_8H_4O_{10}$, 2HO or $\overline{T}$,2HO)

The crystallised tartaric acid of commerce is sufficiently pure for the purpose. It cannot be kept in solution because it soon turns mouldy, and therefore a little of the dry tartaric acid is dissolved only when wanted.

Tartaric acid must dissolve completely in alcohol; if it does not, a salt of lime is probably present, which, however, may

be readily detected by dissolving a little of the acid in water, neutralising with ammonia, and testing with oxalate of ammonia, which will then yield a white precipitate of oxalate of lime.

Tartaric acid is frequently used to prevent the precipitation of several metals; for example, peroxyde of iron, manganese, cobalt, alumina, chromium, &c., by ammonia, or other alkalies, because a double tartrate is then formed which is not decomposed by the alkali. (See Table IX., Observation *p*.)

The concentrated solution of tartaric acid is used as a test for *potash*, and also to distinguish soda from lithia and from ammonia, and even from potash, the acid tartrate of potash being much less soluble in water than the corresponding soda salt.

REACTIONS.

Potash . . .	. . .	In concentrated solutions,
	Crystalline .	precipitate, sparingly soluble in water, soluble in the strong acids and in solutions of carbonate of potash, of caustic potash, soda and ammonia; insoluble in strong alcohol. This precipitate does not immediately appear in solution of sulphate of potash. Violent shaking promotes its formation. (See Table VI., Observations *r*, *s*, *t*.)
Soda . . . / Lithia . . .	*Nothing*.	
Ammonia . .	*White crystalline precipitate*	In concentrated solutions.
	Nothing .	In dilute solutions.
Titanic acid .	*White* . .	precipitate; soluble in acids.
Chromic acid		Reduced to Cr_2O_3, and the liquor becomes
	Greenish .	CO_2 is evolved.

TERCHLORIDE OF GOLD.

($AuCl_3$.)

This reagent is prepared in the following manner:—Take a gold coin, or old articles of gold jewellery, cut into small pieces, introduce them into a flask, and pour upon them an excess of aqua regia; apply heat as long as the mass is seen to be acted upon, adding more aqua regia if necessary. When the whole has dissolved (except, probably, a white sediment, which is *chloride of silver* with which the gold was alloyed), take a small portion of the liquor dilute with water, filter, and test the filtrate with a solution of ferrocyanide of potassium; if this produces *no precipitate*, it is a proof that *no copper* is present, and therefore the liquor may at once be evaporated to dryness by the heat of a steam bath. If, on the contrary, the addition of ferrocyanide of potassium has produced a crimson, or brownish-red precipitate, it is due to the presence of copper which was alloyed with the gold; in that case the whole solution is diluted with water, and filtered, if need be, in order to separate the chloride of silver, and to the filtrate an excess of solution of protosulphate of iron must be added to the gold solution; this addition will produce a *brownish-black* precipitate, which is *pure metallic gold* in a state of extreme division; this brownish-black precipitate is therefore carefully collected on a filter, thoroughly washed, reintroduced in the flask, redissolved in aqua regia, evaporated to dryness by the heat of a steam bath, and the dry residue being dissolved in about 30 parts of water forms a neutral solution of perchloride of gold, which is principally used as a test for *protoxyde of iron*, which produces a dark-brown precipitate of metallic gold; and for *protoxyde of tin*, which gives a purple precipitate (purple of Cassius); yet, in the latter case, the liquor must contain a little peroxyde of tin at the same time (see Table

II.—B., Observation *g*), otherwise no purple precipitate appears.

Terchloride of gold is used also as a test for *formic acid, sulphurous acid,* and *oxalic acid,* and a few other acids which reduce the gold in the metallic state.

TEST-PAPERS.

The papers used for tests are of two kinds; namely, papers dyed with vegetable colouring matter, capable of experiencing a striking and remarkable change of colour when moistened by a free acid, or with certain neutral salts, or with an alkali; and papers imbued with metallic solutions capable of assuming a distinct colour when in contact with certain gases, or with certain metallic solutions, capable of forming coloured precipitates with the solution in which the test-paper has been immersed.

Test-papers, when used for testing gases, should alway be moistened with water before use, but when the substance to be tested with them is an aqueous liquid, this precaution is, of course, superfluous.

The test-papers employed by the analyst are the following:—

Acetate of Lead paper.
Brazil-wood paper.
Ferrocyanide paper.
Ferricyanide paper.
Georgina paper.
Litmus (blue) paper.
„ (red) paper.
„ (grey) paper.
Red Cabbage paper.
Turmeric paper.

All test-papers should be kept in well-stoppered phials.

ACETATE OF LEAD PAPER, or, as it is sometimes called, "*lead paper*," is prepared by immersing filtering paper into a strong solution of basic acetate of lead, hanging it to dry, and then cutting it into strips or pieces of a suitable size. These strips should be kept in glass stoppered bottles

The use of this paper is for detecting the presence of sulphuretted hydrogen. It is exceedingly delicate, the smallest trace of that gas imparts to it a deep brown or black colour.

BRAZIL-WOOD PAPER is prepared by boiling Brazil-wood (called also Pernambuco, Fernambuca, and Sapan-wood) for about an hour in spring water, and dipping strips of white blotting-paper into the decoction.

The paper is chiefly used as a test of the presence of sulphurous acid, by which it is bleached.

FERROCYANIDE PAPER is prepared by immersing filtering paper into a solution of ferrocyanide of potassium, hanging it to dry, and cutting it, when dry, into strips of a suitable size.

This paper is used as a test of the presence of all the metals for which the solution of ferrocyanide of potassium is a reagent. (See Ferrocyanide of Potassium.)

FERRICYANIDE PAPER is prepared like the ferrocyanide paper, by substituting thereto a solution of ferricyanide of potassium.

The principal use of this paper is for the purpose of ascertaining when all the protoxyde of iron in a solution has been converted into peroxyde of iron, because as soon as this is accomplished, the test-paper being touched with the liquid, is no longer rendered blue by it, the persalts of iron having no action upon ferricyanide of potassium. (See Ferricyanide of Potassium.)

GEORGINA PAPER. This paper is prepared by immersing white filtering paper into an infusion of the petals of the purple dahlia (*Georgina purpurea*). The decoction is made

by pouring boiling water on the petals of the flower. After about a quarter of an hour the liquid is filtered, and white bibulous paper is dipped into the filtrate, which should be strong enough to impart a rather deep violet colour to the paper.

This paper is a delicate test for both acids and alkalies, the first turning it *red,* the second *green,* or *yellow* if concentrated.

Litmus Paper (Blue) is prepared by pulverising 1 part of commercial litmus, and digesting it in 6 parts of cold water, filtering, and dividing the blue liquid into two equal portions. To one of the portions add carefully, and one drop at a time, as much sulphuric acid in a very diluted state, as is sufficient to impart to it a slight red colour, this being done, pour the portion so treated into the second portion, which has an intensely blue colour, and stir the whole together; the mixture so obtained is neutral, and by immersing slips of white blotting paper into it, and carefully drying them by hanging them on a stretched piece of thread, an exceedingly sensitive blue test-paper is obtained which should be kept sheltered both from the air and light.

Blue litmus paper is immediately reddened by free acids, but most neutral salts of the more heavy metals produce the same reaction.

Red Litmus Paper is prepared by dipping the strips of blotting paper into the first half of the solution of litmus, reddened by dilute sulphuric acid described in the last paragraph but one. Yet a little more sulphuric acid must be added in the present instance, because the paper must look distinctly red when dry.

Red litmus paper may also be made by adding a drop or two of very dilute sulphuric or hydrochloric acid to a large quantity of water, and immersing blue litmus paper into it, so that it may become red. The paper is then dried as usual.

Red litmus paper is used as a test of the presence of free alkalies, because it is again rendered blue by them; yet the

operator must bear in mind that alkaline earths, the sulphurets and the carbonates of alkalies, restore the blue colour of reddened litmus paper, and the soluble salts of boracic and other weak acids possess also this property.

GREY LITMUS PAPER is prepared by adding as much very dilute sulphuric acid to the solution of blue litmus as to give it a shade between red and blue. The blotting paper dipped in such a liquid is extraordinarily sensitive, and may serve for both acids and alkalies, the smallest trace of which renders it either red or blue. It must, of course, be kept in bottles with sound corks, or glass stoppered, and sheltered from daylight, which otherwise would bleach it after a short time.

RED CABBAGE PAPER.—By macerating red cabbage in hot water, a bluish infusion is obtained with which a test-paper for both acids and alkalies may be prepared in the usual way. Red cabbage paper is rendered *red* by acids, and *bright green* by alkalies. It is not so sensitive as litmus paper.

TURMERIC PAPER is prepared as follows :—Take 1 part of turmeric powder, or of turmeric root, previously pounded in a mortar, and macerate it with 6 parts of spirits of wine, or of wood naphtha; filter; immerse pieces of white filtering paper into the filtrate, hang them to dry, and cut them into strips. The paper has a fine *yellow* colour, and is turned *brown*, or rather *reddish brown*, by alkalies. It is, however, very far from being as sensitive as *red* or *grey litmus* paper, or even *red cabbage* paper. The operator must also recollect that the sulphurets and carbonates of alkalies, and boracic acid have an alkaline reaction upon turmeric acid.

TINCTURE OF GALLS.

Tincture of galls is prepared by digesting two ounces of bruised galls in one pint of proof spirit of wine, and filtering or carefully decanting.

Tincture of galls is used for the same purposes as the aqueous infusion, but the infusion, unless common salt be added, is liable to turn mouldy, which is not the case with the tincture (See Infusion of Galls).

TINCTURE OF LITMUS.

Tincture of litmus is prepared like the tincture of galls, of course, by replacing the galls by litmus.

It is chiefly used to impart a colour to certain solutions, in order to indicate when the point of saturation by an acid is attained, as for example in alkalimetrical operations.

TINCTURE OF TURMERIC.

Tincture of turmeric is prepared like the tincture of galls, and of litmus, merely replacing the litmus or the galls by turmeric powder.

Its principal use is the same as that of litmus, but it is much less sensitive.

TURMERIC. (See Test Papers)

YELLOW PRUSSIATE. (See Ferrocyanide of Potassium.)

WATER.

(HO.)

Distilled water is the most important of all solvents, and the liquid most frequently used and in largest quantity in the laboratory. It is obtained by distilling ordinary spring water from a glass retort, or from a copper or tin vessel of suitable

dimensions. The distillation should be carried at a slow rate in order to prevent impurities from being mechanically carried up with the stream, and for the same reason not more than about three fourths of the charge of water in the still should be distilled. Pure rain water received in the open air, *but not from the roofs of houses,* may generally be substituted, provided, of course, it stands the tests by which the purity of water is ascertained.

A few hundred grains, for example 500 grains, of pure water, being evaporated in a platinum crucible, must not leave any residue whatever.

Pure water should not alter the colour of litmus or other test papers.

It must not be rendered turbid or cloudy by *nitrate of silver,* if it does, a chloride is present.

Nor by *chloride of barium,* which otherwise would indicate a sulphate, probably sulphate of lime.

Nor by *oxalate of ammonia,* otherwise a *salt of lime* is present.

Lime water should not render distilled water turbid, if it does it is a sign of the presence of *carbonic acid,* which, however, may be expelled by boiling.

Pure distilled water is used as a solvent, and for washing; it is used also as a test for the salts of *bismuth,* of *antimony,* and of *tin,* with which it produces a milkiness. (See Table I.—E., Observation *m,* and Table IX., Observation *i.*)

ZINC.

The zinc of commerce is contaminated with *iron, cadmium, lead,* and sometimes other metals which, however, do not materially interfere with its use as a reagent, at least in many cases. It generally contains *arsenic,* the smallest trace of which, of course, renders it perfectly unfit for Marsh's apparatus. Such zinc, however, may be purified by melting

it, and pouring it out into a deep pail of water. The zinc thus granulated is then gathered and placed in alternate layers with one quarter of its weight of saltpetre, into a hessian crucible, taking care to begin with a layer of saltpetre and to terminate with one of zinc. The charged crucible is then heated, and deflagration and fusion having taken place, the scories are removed and the zinc is run into an ingot mould made of chalk, or of plaster of Paris.

The smallest quantity of arsenic in zinc is detected by putting a few pieces of it in a disengagement bottle, and pouring upon it a mixture of pure dilute sulphuric acid, and after the gas has been evolved for some time, bringing the disengagement tube to a red heat by placing a spirit lamp under it, the slightest trace of arsenic will, of course, yield a metallic mirror in the cold part of the disengagement tube.

Metallic zinc is used for the purpose of precipitating certain metals from their acid solutions, such as *copper*, *silver*, *tin*, *tellurium*, *antimony*, *cadmium*, *platinum*, *palladium*, *rhodium*, *iridium*, and *gold*.

Chloride of silver is reduced when treated by zinc and dilute sulphuric acid, the hydrogen evolved combining with the chlorine to form HCl, whilst the silver is set free and may be obtained quite pure by washing. It also reduces *arsenic* acid into metallic arsenic.

Zinc is also used as a means of producing hydrogen gas, and as we said before, for the detection of arsenic by means of Marsh's apparatus. (See Arsenious Acid.)

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